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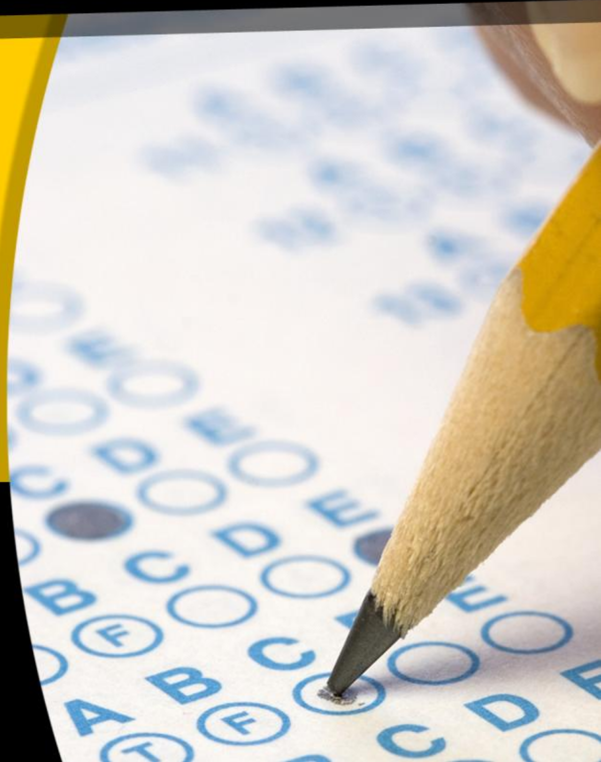

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Written and edited by the Mometrix Exam Secrets Test Prep Team
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Quantitative Chemistry

Stoichiometry

Stoichiometry is the method by which the mass of compounds and how those various masses interact during reactions can be calculated. The law of conservation of mass states that no detectable quantities of mass are lost in a reaction. For example, if a candle is burned in a sealed jar, the mass of the jar and its contents will be the same before burning the candle as after. Stoichiometric calculations, therefore, are important in describing how elements in a compound relate to one another and in predicting the chemical properties of products of reactions. These calculations have practical application in many industrial functions. For instance, they allow manufacturers of ammonia to know how much raw materials (methane, water, and air) are necessary to produce a desired amount of ammonia.

Mole

The mole is a unit of measurement for the number of particles in an amount of matter. A mole contains the same number of particles as is present in exactly 12 grams of carbon-12, or Avogadro's number, 6.0221415×10^{23}. The particles counted must be specified, and may be atoms, molecules, electrons, or other particles. For example, a mol of water may be said to contain either a mol of H2O molecules, or three moles of atoms. The mole is significant because it allows different substances with widely varying density and mass to be measured in consistent and comparable ways, enabling predictions and calculations important to chemical reactions and their products. Molar mass is an important number calculated using the mole; it is the number of grams per mole of a given element or compound.

Variations in usage

Because a mole is a unit of measurement for different types of particles, the term is used differently depending on the units being measured. At its most basic meaning, a "mole of an element" is the amount of that element containing the same number of atoms as 12g of Carbon-12. A "mole of a compound" is the amount of a compound containing Avogadro's number of formula units. A formula unit is the smallest repeating constituent of a compound expressed in its simplest ratio. For example, NaCl is the simplest ratio (1:1) of sodium to chlorine atoms in table salt, so a mole of the compound contains Avogadro's number of NaCl. Similarly, the molar mass of an element is the number of grams per mole of an element, while the molar mass of a compound is the number of grams per mole of a compound.

Calculations

The mole is the number of particles in a given amount of substance and is the foundation of many important calculations. These include the following:

- Changing moles to grams gives the weight in grams of a mole of any substance.

- Changing grams to moles is a calculation giving the number of moles present in a substance when its weight in grams is known.
- When the number of moles is known, the number of particles in a substance can be calculated using Avogadro's number.
- Molar mass calculations give the mass of a given substance per number of moles.

These calculations are important in chemistry because they allow precise measurement of substances present and provide mathematically reliable constants for very large numbers of particles. Considering Avogadro's number of particles (6.0221415×10^{23}) at a time, rather than a single particle, is a powerful element of stoichiometry and all chemical equations.

Mole, molar mass, and Avogadro's number

The amount of a substance that contains as many elementary units (atoms, molecules, etc.) as there are atoms in exactly 12 grams of carbon, 12 is defined in the SI system of units as a mole. The number of units is determined experimentally and the currently accepted value is

1 mole = 6.022 × 1023 units

This number is called Avogadro's number and it is named in honor of Italian scientist Amedeo Avogadro. The molar mass of an element is the mass in grams or kilograms of 1 mole of the element, or 6.022 × 1023 atoms of the element. For all the elements, the molar mass is equal to its atomic mass in atomic mass units.

Volumetric analysis

Volumetric analysis is simply a system of measurement revolving around volume rather than mass. It is used to determine the volume of a given substance which is an unknown portion of another substance; titration is a form of volumetric analysis. Normality is a unit of measurement designed to simplify and speed up the process of volumetric analysis. It is represented in equations as N (but don't confuse it with the chemical symbol for Nitrogen!). The unit is the molarity of a solution expressed in dissociable H or OH ions. For example, a 1 M solution of HCl contains 1 mol/L of reactable H atoms, and so is 1.0 N (signifying normality); it could also be stated as 1 "equivalent" per liter, referring to the value's comparison relative to the reactable H.

Structural formulas

Conventions and uses

Structural formulas represent the bonds within a molecule using simple lines to show which atoms are bound together. The number of lines connecting elements shows the number of pairs of electrons shared between the two atoms. Ranging from quite simple to very complex, structural formulas can be used to represent the ring and chain structures often present in organic molecules. A structural formula for C2H6O2 is shown here. Each line indicates a molecular bond between atoms formed by sharing one pair of electrons. Carbon atoms in compounds often share two pairs (though not in this example), which would be indicated by a double line between the C atoms.

Condensed structural formulas

Because structural formulas can be unwieldy and awkward to write, especially for large molecules, condensed structural formulas were developed as a way to represent the same information in an easier-to-use format. As in structural formulas, condensed structural formulas describe the bonds between atoms and the numbers of atoms present. Unlike structural formulas, however, their condensed cousins do not show all the bonds in a molecule; they instead show the bonding sequence reading the structural formula from left to right. For example, OHC2H2C2H2HO is the condensed structural formula for the compound illustrated in the structural formula definition.

Chemical formulas

Chemical formulas are chemistry's "shorthand." Composed of letters and subscript numbers, they concisely describe the atomic makeup of compounds. The letters are the periodic symbols for each element, and the numbers indicate the number of atoms present of each element. If only one atom is present, no number is needed. A formula is said to be empirical when its subscripts represent the simplest whole number ratio of atoms. For example, H2O, the empirical formula for water, shows that two atoms of Hydrogen are present for every atom of Oxygen. A molecular formula describes the molecular structure in greater detail, but not necessarily the simplest ratio. For example, benzene's empirical formula is CH reflecting a 1:1 ratio, but the molecule actually contains 6 atoms of each element. The molecular formula is thus C6H6. Any compound can be represented with a formula using these conventions. Formulas are pronounced by stating each letter and number individually.

Solutions

Types

Polar and nonpolar solutes and solvents combine in four different ways to form solutions with varying properties. Here are the four possible combinations: polar solute + non polar solvent; polar solute + polar solvent; nonpolar solute + nonpolar solvent; nonpolar solute + polar solvent. Following the "like dissolves like rule," nonpolar solvents dissolve nonpolar solutes in part because little energy is required to separate the weak London forces governing the interaction of nonpolar substances. Polar substances interact through both London forces and dipole-dipole forces, so they require much larger amounts of energy to separate, or to enter into solution. However, for the same reason, they also make very stable solutions, requiring energy to return to separate solute and solvent states. Solutions of unalike molecules require a great deal of energy (they are endothermic) because the dissimilar molecules repel one another, making solutions unlikely to form.

Ideal solutions

An ideal solution is one in which all molecules, whether solvent or solute, interact in the same way. That is, all attractive forces between molecules in such a solution are equal, so the net enthalpy, or change in energy or heat, is zero. Ideal solutions are useful as calibrators for calculated values in colligative properties; because the enthalpy is zero, theoretical calculations of properties such as osmotic pressure or changes in boiling point can be experimentally confirmed and/or adjusted by comparing them to the known values of ideal solutions. Another way to think of them is that they always follow Raoult's Law exactly, but they are largely mathematical substances—solutions behave in ideal ways only in very dilute forms and are easily subject to change.

Solvation

Solvation is the technical name for the action of dissolving. When a solute mixes uniformly with a solvent so that the two interact on molecular level, this is solvation. Solvation occurs when intermolecular attractions between solvent and solute are greater than the forces holding either the solute or solvent together. The process can be either exothermic or endothermic. The process of dissolving table salt in water is a common example. $H^+{}_2O^-$ and Na^+Cl^- molecules are highly polar due to their chemical properties—specific regions of the molecules have electrical charges. When NaCl is added to water, the polar H_2O molecules are immediately attracted to the opposite poles of the NaCl molecules. Because this attraction is stronger than those holding the NaCl in a crystal lattice, the Na and Cl ions break apart and move into solution—the salt is "dissolved."

Colligative properties and the "like dissolves like" rule

Colligative properties are physical properties of a substance that are influenced by the concentration of a solution. For example, pure methane, C_6H_6, boils at a given temperature under known atmospheric pressure. If a solute is added to the methane, the boiling point will no longer remain constant. The colligative properties are: vapor pressure reduction, boiling point elevation, freezing point depression, and osmotic pressure.

Like dissolves like refers to the fact that a solvent must be similar in polarity to a solute if the solute is going to dissolve. Salt dissolves easily in water because both are polar substances. Water does not dissolve oil because the lipids and other hydrocarbons of oily substances are not polar. Gasoline, a nonpolar solvent, dissolves nonpolar molecules such as oil and grease.

Measurement of concentration

Solutions can be measured in everyday applications in terms of percentages. For example, a label on a vinegar bottle that says 5% solution means that 5% of the total solution is acetic acid. Concentrations are also measured by mass; HCl purchased for laboratory uses, for example, is usually 37% by mass, meaning that for every 100g of solution, 37g of HCl are present. Parts per million is another term of measurement for very dilute solutions. For more precise concentration measurements, chemists calculate mole fractions, molarity, and molality. Molarity, represented in equations by M, is the number of moles of solute per liter of solution. Molality, represented in equations by m, is the number of moles of solute per kilogram of solvent.

$$M = \frac{moles\ A}{liters\ of\ solution}$$

$$m = \frac{moles\ A}{kg\ of\ solution}$$

Henry's Law

Henry's law defines a constant used in solution equations. It states that the amount of a gas dissolved in a liquid at given temperatures and concentrations corresponds directly to the partial pressure of the gas over the liquid. Or in other words, it states that the volume of a gas that is soluble is proportional to the partial pressure of that gas above the solution in a closed system. Partial pressure is defined as the pressure of a gas if it occupied a container all on its own. It is calculated as $P=kC$, where k is the Henry's law constant, C is the concentration of the gas in moles per liter, and P is the partial pressure. Henry's law is one of the gas laws describing the thermodynamic interaction of temperature, pressure, and volume in gases. The law establishes constants at 25° C.; because pressures change as temperatures change, temperature must be considered when using Henry's Law.

Non water-based inorganic solvents

For much of chemistry's early history, water was the only known solvent. Others gradually came to be used, but usually in aqueous solution. Later developments demanded non-aqueous solvents to study particular reactions that cannot occur in water. Many non aqueous solvents are organic compounds, but some common ones, such as liquid ammonia and sulfur dioxide, are not. In addition to research uses, these types of solvents are significant in many industrial processes. Acids and bases often behave differently in such solvents than in water or water-based solutions. For this reason, it is important to consider the active ions when measuring acidity and basicity of substances in inorganic non aqueous solutions.

Factors affecting solubility

The most obvious factors which influence solubility are temperature and pressure. An increase in solvent temperature increases the solubility of the solid solute—the higher the temperature, the more solute can be dissolved in solution. The impact on temperature is not uniform across substances, and there are a few examples of solvents whose solubility *decreases* with increase in temperature, but these are rare. Gaseous solutes, however, become less soluble at higher temperatures. Pressure impacts solubility only when solvents are in the gaseous state. Much like a system in equilibrium, gaseous solvents move out of solution and back into it under sufficient atmospheric pressure. If atmospheric pressure is not strong enough to force molecules back into solution, the gas escapes into the air—all gaseous solutes will eventually leave solution in an open container.

Conductivity as a physical property of solutions

Pure solvents containing no free ions have very low electrical conductivity; most water has high conductivity because it is not pure H2O—it is a solution containing salts and other minerals whose ions conduct the electrical current. Like the colligative properties, conductivity is influenced by the concentration of solute within a solution. Therefore, water with a lot of table salt (Na^+Cl^-) dissolved in it (a solution) has much greater electrical conductivity than distilled water, which has been processed to remove most substances other than water molecules. However, even the purest water produced in a lab has some conductivity because of the slight tendency (approximately 1 in 50 million) of its H atoms to dissociate and form ions.

Energy in solutions

Although solutions may appear inert and static, they require constant and significant exchanges of and fluctuations in energy. Energy is required at each step of the solvation process, which may be simplified as follows: 1) Solute molecules held together by intermolecular forces must separate; 2) Solvent molecules must also separate to create "room" for the solute; 3) Solvent and solute molecules interact and bond through intermolecular forces. Steps one and two are endothermic, taking energy from outside the system. Step three is slightly exothermic—the energy released is very small. The sum of the changes in heat energy, or enthalpy, for each step is called the heat of solution ($\Delta H^\circ_{sln.}$), and may be either negative (net exothermic) or positive (net endothermic). When it is negative, the intermolecular attraction between molecules in solution is stronger than the attraction between separate solute molecules.

Precipitation reactions

In terms of solutions, precipitation occurs when a solid precipitates out of solution. In typical precipitation reactions, ionic solutes first exist in equilibrium in separate solutions. When the solutions are combined, the ions combine into an insoluble solid which then drops out of solution. The reactions are often quite dramatic, producing bright colors from the mixture of two clear solutions. Consider the following example of a precipitation reaction involving silver nitrate and table salt in separate solutions: $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$. The solid, silver chloride, is insoluble in water and forms because the intermolecular attraction between Ag^+ and Cl^- is stronger than that which holds NaCl together. The solid then drops out of solution, appearing as a gray cloud spreading throughout the solution. The reaction may also be expressed as a net ionic equation: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$.

Vapor-pressure reduction and Raoult's Law

When a solute is added to a solvent, the vapor pressure of the resulting solution will be lower than that of the pure solvent. This reduction in vapor pressure is caused by molecules and larger particles of solvent at the surface of the solution, where particles of solvent can escape, i.e., where the solvent can be vaporized. Because of the presence of solute particles, fewer particles of liquid solvent can escape into vapor—the solute particles literally get in the way. Raoult's law, named for the first chemist to investigate the mathematical relationship of vapor pressure to solute concentration, states that the vapor pressure of solution is equal to the vapor pressure of pure solvent times the solute concentration expressed in mole fractions. The equation is as follows: $P_A=P°_A X_A$, where P° is the vapor pressure of the solvent and X is the mole fraction of the solvent.

Boiling point elevation and electrolytic solutes

Substances boil when their vapor pressure—the tendency of liquids or solids to enter gaseous states—equals the atmospheric pressure pressing upon them. When electrolytic solutes—substances that separate in solution into ions—are present in a solution, the boiling point increases because the vapor pressure has been reduced per Raoult's law. Higher temperatures are therefore required to produce sufficient energy to allow particles to escape from solution to vapor. Non electrolytic solutes do not have the same influence on boiling point because they release fewer particles into solution upon solvation. Table sugar, for instance, releases smaller particles into solution when it dissolves, with the smallest possible unit being a single sucrose molecule. However, electrolytic solutes release more particles because molecules are broken apart into individual ions. This number or particles must be considered when calculating the effect of solutes on boiling point.

Freezing point depression and electrolytic solutes

In converse fashion, the addition of electrolytic solutes results in a solution with a lower freezing point than the pure solvent. The additional particles make it more difficult for the liquid solution's molecules to align themselves in a crystal lattice; they require a lower temperature to do so. Water containing large amounts of dissolved salt can thus be cooled below its normal freezing point of 0° C. As with boiling point elevation, the mathematical calculation of freezing point depression depends on the number of particles released during solvation. Electrolytic solutes break apart into individual ions, while non electrolytic solutes can only be reduced to individual molecules. Most organic solutes are non electrolytic so have a much smaller effect on colligative properties than inorganic compounds such as salts.

Osmotic pressure

Osmotic pressure is the pressure exerted by a solvent passing through a membrane in osmosis. It is an important measurement used to predict the concentration of solute required to establish equilibrium on each side of the membrane. Imagine two different concentrations of solution are separated in a closed system by a permeable membrane; water can pass through the membrane, but solute particles do so at a much lower rate. Water will flow from the side with lower concentration to the side with higher concentration, but the greater number of solute particles on the higher concentration side reduce the number of solute particles that are able to pass to the lower side. Because water flows freely through the membrane and solute particles less so, the osmotic pressure on the high concentration side increases, causing water to flow back through the membrane to the lower side until equilibrium is established.

Dielectric constant

The dielectric constant is a measurement of a solvent's ability to dissolve polar substances; it is symbolized in equations by ε. The higher the dielectric constant, the more likely a solvent is to dissolve polar solutes; such solvents are usually highly polar themselves. The value is derived experimentally and is more accurately a reflection of the interaction between solvent and solute rather than a property inherent in the solvent. Water's ε is very high, 78.5, while hexane has a very low dielectric constant of 1.89. Hexane will therefore not dissolve ionic compounds. Other common solvents and their ε are as follows: benzene, 2; diethyl ether, 4; tetrahydrofuran, 7; acetone, 21; ethanol, 24; methanol, 33; dimethyl sulfoxide, 49; and water, 78.5.

Atomic Structure

Matter

Matter is defined as anything that occupies space and has mass. Categories of matter include atoms, elements, molecules, compounds, substances and mixtures.

- An atom is the basic unit of an element that can enter into a chemical reaction.
- An element is a substance that cannot be separated into simpler substances by chemical means.
- A molecule is the smallest division of a compound that can exist in a natural state.
- A compound is a substance composed of atoms of two or more elements chemically united in fixed proportions.
- A substance is a form of matter that has a definite or constant composition and distinct properties.
- A mixture is a combination of two or more substances in which the substances retain their unique identities.

Physical and chemical properties of matter

If a property of a substance can be observed and measured without change it is a physical property. Density, melting point and boiling point are examples of physical properties. If a chemical change must be carried out in order to observe and measure a property, then the property is a chemical property. For example, when hydrogen gas is burned in oxygen, it forms water. This is a chemical property of hydrogen because after burning a different chemical substance – water – is all that remains. The hydrogen cannot be recovered from the water by means of a physical change such as freezing or boiling.

States of matter

The three states in which matter can exist are as a solid, liquid or gas. They differ from each other in the distance of separation between the individual molecules. In a solid, the molecules are very close to each other with little range of motion. Molecules in a liquid, while close, have a greater degree of freedom than molecules in a solid and can move around each other. In a gas, the molecules are separated by distances that are very large in comparison to the size of the molecules. The three states of matter can be interchanged by the addition or removal of heat. For example, when a solid is heated to its melting point, it will form a liquid. Upon further heating to its boiling point it will form a gas.

<u>Characteristic properties</u>
The table below outlines the characteristic properties of the three states of matter:
<u>Gas</u>
Volume/shape: Assumes volume and shape of its container
Density: Low
Compressibility: High

Molecular Motion: Very free motion

Liquid

Volume/shape: Volume remains constant but it assumes shape of its container

Density: High

Compressibility: Slightly

Molecular Motion: Move past each other freely

Solid:

Volume/shape: Definite volume and shape

Density: High

Compressibility: Incompressible

Molecular Motion: Vibrates around fixed positions

Properties of liquid

Liquids have definite volumes, however, they are unable to change shape by flowing. Liquids and solids have similarities in that in each the particles touch but they can move around. This means the density of liquids will be near that of a solid. Because liquid molecules can move, they assume the shape of their container. Specific liquid properties include viscosity, or the resistance of a liquid to flow. Surface tension, when there is an attraction between the molecules of the liquid that causes the surface to act as a thin and elastic film that is under tension, is another property. This property is what causes the formation of water into spherical drops. Vapor pressure is the pressure a solid or liquid exerts when at equilibrium with its vapor at a particular temperature. Boiling point is when atmospheric pressure equals the boiling point.

Density of water

In the solid state, water is less dense than in the liquid state. This can be observed quite simply by noting that an ice cube floats at the surface of a glass of water. Were this not the case, ice would not form on the surface of lakes and rivers in those regions of the world where the climate produces temperatures below the freezing point. If water behaved as other substances, lakes and rivers would freeze from the bottom up and be detrimental to many forms of aquatic life. The lower density of ice occurs because of a combination of the unique structure of the water molecule and hydrogen bonding. In the case of ice, each oxygen atom is bound to four hydrogen atoms, two covalently and two by hydrogen bonds. This forms an ordered, roughly tetrahedral structure that prevents the molecules from getting close to each other. As such, there are empty spaces in the structure that account for the low density of ice.

Crystalline solids

The atoms of solids are closely packed and are not compressible. Since all solids have some thermal energy, the atoms vibrate. The movement is very small and rapid and cannot be seen under normal circumstances. The four types of crystalline solids have a specific melting point and contain bonds that are ionic. Sodium chloride is an example. Covalent solids appear as one large molecule that is made up of a virtually limitless number of covalent bonds. Graphite is an example.

Molecular solids are represented as repetitive units that consist of molecules such as ice. Metallic solids are also repeating units. They contain metal atoms. The valence electrons in metals can jump from one atom to another.

Amorphous solids

Amorphous solids have no definite melting point nor do they have regular repeating units. These are solids with no long-range order in positioning of the atoms. This is unlike atoms in crystalline solids. Window glass is an example of amorphous solids. Many polymers such as polystyrene are also amorphous. These solids may exist in two separate states. One is a “rubbery” state. The other is a “glassy” state. The transition temperature between the rubbery and glassy states is known as their glass transition temperature (Tg). More specifically, it is a temperature at which a loss of flexibility becomes lost in adhesives. Adhesives will become hard and inflexible. Adhesives which reach the Tg may fail if they are flexed. The actual temperature widely varies from 105 degrees C to less than 0 degrees C.

Plasma

Plasma is an ionized gas. It is a gas in which enough energy is provided to free the electrons from atoms and molecules. This allows both ions and electrons to coexist. Plasma can be viewed as a cloud of protons, neutrons and electrons in which the electrons detach from their molecules and atoms. This enables the plasma to act as a whole rather than a conglomerate of atoms. More than 99 percent of the visible universe is made up of plasma, most of which is visible, making it the most common state of matter. Plasma is naturally-occurring. It is what makes up the sun, the core of stars and is also found in quasars.

Thermoplastic

Thermoplastic material are materials that are plastic, melts to a liquid when it is heated and freezes into a glassy state whenever it is sufficiently cooled. Most thermoplastics are polymers with high molecular weights whose chains are associated through weak van der Waals forces such as polyethylene, stronger dipole-dipole interaction and hydrogen bonding such as nylon as well as the stacking of aromatic rings such as with polystyrene. Thermoplastic polymers are unlike "thermosetting" polymers such as vulcanized rubber which cure through added energy to a stronger form and can never be remelted or remolded once they are formed. Many thermoplastic materials are so-called "addition polymers" which are vinyl chain-growth polymers such as polypropylene.

Fifth state of matter

A condition in which atoms collapse into a single quantum state is called “Bose condensation” or “Bose-Einstein condensation.” It is known as a possible fifth state of matter. Bose-Einstein condensate has been thought of by scientists as plasma's opposite. It is found at extremely low temperatures that are almost to the point where atoms do not move. The Bose-Einstein condensate consists of a phase of gaseous superfluid that has been formed by atoms and has been cooled to temperatures near absolute zero. The phenomenon was forecasted in the 1920s by Satyendra Bose and Albert Einstein.

Subatomic particles

The three subatomic particles are the proton, electron and neutron. Protons are the positively charged particles in the nucleus, with a mass of 1.67252 x 10-24 g. This is approximately 1840 times the mass of the oppositely charged electron. Neutrons are also present in the nucleus, are electrically neutral and have a mass slightly greater than that of protons. The table below summarizes the three particles:

Particle	Mass (g)	Charge
Electron	9.1095 × 10-28	-1
Proton	1.67252 × 10-24	+1
Neutron	1.67497 ×	0

Shells

Interactions between electrons are what cause the chemical behavior of atoms. Electrons of atoms remain with certain, predictable configurations of electrons. Quantum mechanics determine the configuration of electrons in what is the atom's electric potential. The principal quantum number determines the particular electron shells with energy levels that are distinct. Usually the higher a shell's energy level is, the farther it is away from the nucleus. Valence electrons in the outermost shell have the most influence on chemical behavior. Core electrons, those not in the outer shell, have a role that usually is along the lines of a secondary effect because of screening the positive charge in the atomic nucleus.

Nucleon properties

Nucleon is the name for the two baryons (subatomic particles), the neutron and the proton. They are constituents of atomic nuclei and were thought to be elementary particles as late as the 1960s, their interaction then defined strong interactions. They now are known to be composite particles made of gluons and quarks. Constituent protons and neutrons of the atomic nucleus are known as nucleons. They are held together by a strong nuclear force in the nucleus. Nuclei transformation can play a part in radioactivity. Nuclear transformation also occurs in nuclear reactions. In nuclear fusion, two light nuclei merge into a single nucleus that is heavier. In nuclear fission, a single large nucleus divides into two or more smaller nuclei.

Electron mobility

Electrons are accelerated in an electric field, E, in an opposite direction to that field due to its negative charge. The force that acts on the electron is -eE, where e is the charge. This force causes a constant acceleration in that in the absence of obstacles in a vacuum, the electron goes

continuously faster in an electric field. In solids, a different situation occurs; the electrons scatter due to collisions with atoms and vacancies that drastically change the direction in which they move. So, electrons move randomly. But they also move with a net drift that is in the opposite direction to the electric field. This drift velocity is constant and is equal to the electrical field multiplied by a constant known as the mobility, m.
Va = me E

Electron configuration

Electron configuration, in atomic physics, is how electrons in an atom, molecule, or other body are arranged. Expressly, it is the electron placement into atomic, molecular, or other forms of electron orbitals. The state of an atom's electrons is given by four quantum numbers. Three integers are properties of the atomic orbital in which it is located. No two electrons in one atom may possess the same set of these four quantum numbers as in the Pauli exclusion principle. A standard notation is used to describe atomic electron configurations; a subshell is written in the form nxe, where n is the shell number, x is the subshell, and e is the number of electrons present in the shell.

Unit cells

A unit cell is the smallest structure that repeats itself via translation through a crystal. These are symmetrical units that have hard spheres. Most common among unit cells are the faced-centered cubic, hexagonal close-packed and body-centered cubic. Properties that are important in unit cells are:

- The type of atoms and their radii R.
- Cell dimensions -- side a in cubic cells, side of base a and height c in hexagonal cells, in terms of R.
- n, number of atoms per unit cell. A fraction of the atom, 1/m, is counted for an atom that is shared with m adjacent unit cells.
- CN, coordination number, the number of closest neighbors to which atoms bond.
- APF, the atomic packing factor. This is the fraction of the volume of the cell that is occupied by hard spheres. APF = Sum of atomic volumes/volumes of cell.

Plum pudding model

J.J. Thomson suggested the arrangement of protons and electrons within an atom were a "plum pudding" model. Thomson, who discovered the electron before the proton or neutron, hypothesized that an atom has electrons that are encompassed by a blob of positive charges that surround it like pudding. The electrons were believed to be positioned uniformly inside the atom and was also said to sometimes possess a cloud of positive charge. By studying electricity that passed through gases, Thomson found that all atoms have certain particles, called electrons, and that they carry a negative charge. Since atoms are electrically neutral there must be positive charges somewhere in the atom to balance the electrically negative ones. This model was later disproved.

Geiger-Marsden experiment

An experiment by Hans Geiger and Ernest Marsden (under the supervision of Ernest Rutherford) at the University of Manchester led to disproving Thomson's plum pudding model of the atom. Geiger and Marsden measured the deflection of alpha particles that were directed normally onto a sheet of thin gold foil. Thomson's model would have seen the alpha particles all being deflected by at most a few degrees. But they saw a very small percentage of particles were deflected through angles that were much larger than 90 degrees. This led to additional findings later by Rutherford that suggested the atom had a very small positive charge which was able to repel the alpha particles if they were close enough.

Size and speed of atoms

Electron clouds do not have a sharp cutoff, so the size of an atom is hard to define. When atoms form crystal lattices, the distance between the centers of the adjacent atoms can be found by X-ray diffraction. This gives an estimate of the size of an atom. The radius at which the valence shells are most likely to be seen can be used for an atom. Temperature of an atom collection is measured by the average energy of motion of the atoms. Kinetic energy of the particles in the system increase as temperature increases and speeds up the motion. At room temperature, atoms that form gases in the air will move at speed of about 500 m/s or 1,100 mph.

Distillation and fractional distillation

Distillation is the process by which different components of a mixture may be separated. Because substances in the mixture have different boiling points, some are vaporized as the temperature rises, leaving others in the liquid state. The vapors may be trapped and condensed to return them to a liquid or solid state. Fractional distillation is similar but uses a different experimental apparatus. A tower is attached to a flask containing the mixture to be distilled. As vapors rise, they are trapped and condensed, returning to the flask to repeat the process. On an industrial scale, the refinement of crude oil is based on fractional distillation; different substances such as gasoline, kerosene, and diesel are produced as rising vapors cool and condense at different temperatures.

Vaporization

The evaporation of part of a liquid will see cooling of the liquid left behind. That is because it has to extract the needed heat of vaporization from the liquid so it can change into a gaseous state. It is an important mechanism in heat transfer. The human body is cooled when it has ambient temperatures that are above the normal body temperature. Evaporation is an effective method of cooling because of the large heat of vaporization of water.

The human body cools off by perspiring to give energy, even though the temperature around the body is a higher than that of the body. The cooling rate for a liquid that is below boiling will see of heat of a vaporization change with temperature and the rate of evaporation will depend upon the relative humidity and ambient temperature.

BLEVE

A boiling liquid expanding vapor explosion (BLEVE) is a danger in many railroad accidents in which flammable chemicals are involved. It is an explosion that happens when a vessel, such as a railroad tank car, fails and inside the vessel is a liquid at a temperature well above its boiling point at normal atmospheric pressure. With such a failure, boiling can occur instantly depending on the particular boiling point of the liquid. This causes the vaporization of a large amount of the liquid to occur. If the liquid's temperature is higher than the superheat limit, then boiling of an instantaneous nature happens. If the temperature is below that limit, the energy for blast and fragmentation generation is released. This mainly comes from vapor expanding in the space above the liquid. In either instance, if the container is burning the metal is heated and this causes a loss of mechanical strength.

Vacancy and interstitials

A vacancy is a position on a lattice that is vacant due to a missing atom. It comes about upon formation of the solid. Other vacancies can be made and also take place naturally through thermal vibrations. An interstitial is an atom occupying a place outside the normal position on a lattice. It might be an impurity atom or a self-interstitial atom, which is the same as other. With vacancy and interstitials, a change in coordination of atoms around the defect takes place. This means forces are not balanced for other atoms in the solid in the same way. This results in a lattice distortion around the defect. The number of vacancies from thermal agitation follows:

$$NV = NA \times \exp(-QV/kT)$$

Where NV = total number of atoms in the solid, QV is the amount of energy needed to form vacancies, k is Boltzmann constant and T is Kelvin temperature.

Impurities in solids

All solids that are real are impure. Impurities are often added to materials in order to make the properties better. For example, carbon added in small amounts to iron produces steel. Boron impurities added to silicon causes a drastic change in its electrical properties. Solid solutions are made of a host, a solvent or matrix, which dissolves the solute. The ability to dissolve is solubility. Solid solutions have certain characteristics such as being homogeneous, maintaining crystal structure, containing randomly dispersed impurities which are either substituted or interstitial. For high solubility there must be a atom size that is similar, have similar crystal structure, possess electronegativity that is similar and have a similar valence.

Energy band structures

Atoms coming together and forming a solid have valence electrons interacting due to Coulomb forces. They also feel the electric field their own nucleus produces and that of other atoms as well. Two specific quantum mechanical effects also occur; first, through Heisenberg's uncertainty principle, electrons constrained to a small volume see its energy rise, or promotion; second, because of the Pauli exclusion principle there are limits to the number of electrons having the same property. Accordingly, the valence electrons of atoms make valence bands that are wide when forming a solid. These bands are kept apart by gaps where electrons cannot exist. In insulators and semiconductors, there are filled valence bands. No more electrons can be added, as per Pauli's principle. Electrical conduction needs electrons that can gain energy in an electric field. Since it would imply the electrons are promoted into the forbidden band gap, this is not possible.

Mössbauer effect

Rudolf Mössbauer discovered the physical phenomenon later named the Mössbauer effect in 1957. It describes the resonant and recoil free emission and absorption of gamma rays by atoms bound in a form that is solid. Mössbauer observed resonance in solid iridium, which led to the question of why gamma ray resonance was not possible in gases, but were possible in solids. Mössbauer postulated that for atoms bound into a solid, under certain conditions, a fraction of the nuclear events could take place virtually without recoil. His observation was attributed to this recoil-free fraction of nuclear events. Emitted gamma rays correspond to nuclear transition energy minus energy lost to recoil during atom emission. If the recoil energy is small then the gamma ray energy still corresponds to the nuclear transition energy and the gamma ray can be absorbed by a second atom of the type of the first. This is called resonance.

Mass balance

Mass balance is the enumeration of material that enters and leaves a system. The balance works on the conservation of mass principle in that matter cannot be created or disappear. Mass that enters a system must either leave or accumulate within the system:

$$\text{IN} = \text{OUT} + \text{ACC}$$

where IN means what is entering the system, OUT represents what leaves the system and ACC means accumulation within the system. Mass balances often are developed for total mass that crosses a system's boundaries but they may also focus on a particular element or chemical compound. When a mass balance is written for a specific compound instead of for total mass of the system, a production term (PROD) comes into play so that:

$$\text{IN} + \text{PROD} = \text{OUT} + \text{ACC}$$

The production term may then describe chemical reaction rates or other factors.

PROD might be positive or negative just as is the case for ACC.

Ionization and solubilizing groups

Ionization is a term used in various ways in chemistry, all revolving around the behavior and properties of charged particles, or ions. For solutions, ionization refers to dissolved polar solutes; their separated ions, e.g. Na^+ and Cl^-, are separated from each other by molecules of solvent. This solution ionization is the basis for solution chemistry. A solubilizing group is a group or portion of a molecule that acts to increase the solubility of the molecule to which it is attached. Such groups usually act by polarizing the molecule, or specific regions of it, allowing it to interact more freely with solvent molecules of similar polarity. This concept is particularly important in organic chemistry because many vital hydrocarbon chains are not water soluble. They become water soluble, however, when a functional group such as carboxyl is attached.

Molecular structure of compounds

The molecular structures of salts, oxides, peroxides, acids, and bases are:

- Salts: In binary compounds, formed of any metal and any non-metal. In ternary compounds, formed of a metal and a complex ion.
- Oxides: any metal and oxygen.
- Peroxides: oxides containing one more oxygen atom than normal valence rules would predict.
- Acids: In binary compounds, hydrogen and any non-metal in solution; in ternary compounds, hydrogen and a complex ion in solution. In reactions, acids can gain or bond with an H ion.
- Bases: A ternary compound made of any metal and the complex ion OH-. Bases can lose H ions in reactions. Acids and bases together form a class of reactions called neutralization reactions. These reactions are important in many industrial and biological processes.

Properties of metals

Some of the characteristic properties of metals are as follows:

- Most metals are malleable and can be pounded into thin sheets; most are also ductile which means they can be drawn out into thin wires.
- All metals are solid at room temperature with the exception of Mercury, which is a liquid at room temperature.
- Metals usually have low ionization energies and they normally lose electrons in chemical reactions. Compounds consisting of metals <u>and</u> non-metals are usually ionic.
- Most metal oxides are basic oxides and those that will dissolve in water form metal hydroxides:
- Metal oxide + water -> metal hydroxide
- Na2O(s) + H2O(l) -> 2NaOH(aq)
- CaO(s) + H2O(l) -> Ca(OH)2(aq)

- Metal oxides show their basic chemical nature by reacting with acids and forming salts and water.
- Metal oxide + acid -> salt + water
- MgO(s) + HCl(aq) -> MgCl2(aq) + H2O(l)
- NiO(s) + H2SO4(aq) -> NiSO4(aq) + H2O(l)

Mass

Mass is a property of a physical object quantifying the amount of matter it contains. The mass of an object, unlike weight, stays the same regardless of its physical location. Inertial mass is a measure of the resistance of an object to changing its state of motion when a force is applied. An object with small inertial mass changes its motion more readily and an object that has a large internal mass does not change as readily. Passive gravitational mass refers to the measure of the strength of an object's interacting with the gravitational field. Within a gravitational field an object with a smaller passive gravitational mass will experience a smaller force than an object possessing a larger passive gravitational mass. This force is called the weight of the particular object. Active gravitational mass measures the strength of the gravitational field because of a particular object.

Atomic mass scale

Atoms of different elements all have different masses. In early research, it was found that separating 100 grams of water into its basic elements resulted in 11.1 grams of hydrogen and 88.9 grams of oxygen. Once it was later discovered that water has two atoms of hydrogen for every atom of oxygen, which implied that an oxygen atom must weigh about 16 times as much as a hydrogen atom. To have a base unit of measurement, scientists developed the atomic mass unit (amu), and standardized it against the 12C isotope of carbon (amu = 12). Using this unit of measurement, the mass of the hydrogen atom (1H) is 1.0080 amu and the mass of an oxygen atom (16O) is 15.995 amu. Once the masses of atoms were determined, the amu could be assigned an actual value:

Average atomic mass

- 1 amu = 1.66054 x 10-24 grams

or conversely:

- 1 gram = 6.02214 x 1023 amu

The average atomic mass of each element is also referred to as its atomic weight, which is listed in the periodic table.

Molecular weight

Molecular weight equals the sum of atomic weights of the atoms in the molecule. To calculate molecular weight, find the average weight or mass of the elements listed in the periodic table. The mass can be determined by knowing their particular molecular formula. For example, the weight of Methane, or CH4, can be determined as such: MW{Methane} = (1 carbon atom per molecule)*(12.011 grams per mole for carbon atoms) + (4 atoms of hydrogen per molecule)*(1.00797 grams per mole for hydrogen atoms) = 16.0429 grams per mole methane

In order to calculate the weight in grams of one mole of molecules of the compound, known as the isotopically averaged molecular mass, use the average atomic weight (in g/mol) of each of the elements in the molecular formula multiplied by the number of times each element is found in each molecule.

Chemical and atomic number interaction of atoms

Mass is concentrated in a nucleus, which is 1/10,000 the overall size of the atom, and contains neutrons and protons. Most of the volume of an atom is filled by the orbiting electrons. An atom's chemical interactions are determined by its cloud of electrons. The outermost electrons determine the basic chemical properties of the atom. Atoms with the same total number of electrons are said to be chemically "identical." Atoms attract or repel electrons until they have a balanced charge. The number of protons in the nucleus equals the number of electrons that are in a neutral atom. This also determines the chemical identity of the atom:
Number of protons = Atomic number

Quantum numbers

Schrödinger's model allowed for an electron to occupy a three-dimensional space. It is thus required that three coordinates, or quantum numbers, describe those orbitals in which electrons are found. The coordinates are the principal (n), angular (l), and magnetic (m) quantum numbers. These numbers describe size, shape and orientation in space of an atom's orbitals. The principal quantum number (n) describes the orbital's size. For example, those orbitals for which n = 2 are larger than those for which n = 1. The angular number (l), describes the shape of the orbital. Orbitals with spherical shapes are (l = 0), (l = 1) for polar, or (l =2) for cloverleaf. The (m) describes the orbitals' orientation.

Stereochemistry

Stereochemistry is the branch of chemistry which studies and describes the three-dimensional spatial arrangement of atoms and molecules and is an extension of molecular geometry. It is particularly concerned with stereoisomers—versions of a given compound having the same chemical compositions but different configurations. Configuration is the arrangement of atoms in space, within a molecule, a variable having important effects in biological processes. Even a relatively simple molecule, such as a C tetrahedron bonded with four groups of atoms, has a large number of possible variations in configuration. The configuration of molecules has important effects on its reactivity and physical and biological processes. Thalidomide provides a dramatic example of the importance of isomers. One isomer of the drug prevents morning sickness while the other was found to cause birth defects. The only difference between the two versions is the molecular configuration.

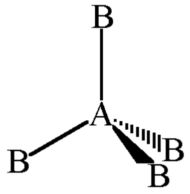
Absolute configuration

Absolute configuration of a molecule is the exact position and orientation of its constituent atoms and groups of atoms. Because molecules are three dimensional, it is important for scientists to know "which way is up" when a given molecule is under consideration, especially in biological processes when the molecular configuration can have enormous consequences. As with nomenclature, standards have been developed which specify the positions of all atoms in the molecule. Absolute configuration can thus be discovered by comparing it to a reference compound whose configuration is known. Furthermore, it can be described by assigning each of the substituents atoms or groups a priority number based on atomic number; the lowest priority atom or group is then said to be "the bottom" of the molecule. Absolute configuration is foundational for biochemistry because certain organic molecules may only "fit" into biological receptors if they are the correct shape and configuration.

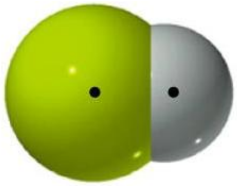
Models and lines for envisioning molecules

Atoms and molecules are complex three dimensional objects in motion. Although they and their bonding properties and structures can be represented with two dimensional drawings such as dot structures and structural formulas, the best models are those which represent the third dimension as well. Because this is difficult to do on two dimensional surfaces such as pages of a book, the following conventions represent three dimensions in molecular drawings:

- Solid lines represent chemical bonds on the plane of the page.
- Wedge-shaped lines represent bonds projecting forward out of the page.
- Dashed lines represent bonds projecting backward out of the page.



Three dimensionality can also be shown by pictures of ball-and-stick molecular models and by "space-filling" models pictured as overlapping spheres.



Lewis dot structure

Lewis structures, also called Lewis electron dot formulas, are symbols in which an atom's valence electrons are represented as dots. The structures consist of the atom's periodic symbol surrounded by dots equaling the number of its valence electrons. Hydrogen has one valence electron, so its electron dot formula is H•, while Helium's is He: with its two valence electrons. Noble gases, with eight valence electrons, are represented with one pair of dots on each side of the element’s symbol. They are useful for illustrating the transfer of electrons in reactions, particularly in those forming ionic compounds. They are also another way to represent the bonds and structure of simple molecules just as a structural formula would do.

Although the Lewis structure is a much simplified representation of the electron configuration of an atom, it is useful for thinking about molecular geometry. The dot structures predict the rough shape of the molecule through covalent bonds that will form by the location of their unpaired electrons. The example here shows the dot structures for H and Fl and the linear (the nuclei are aligned in a singular dimension) shape of an HF molecule. The unpaired valence electrons are represented by a single dot and are also the sites of the covalent bond joining the atoms. The bonded molecule is represented by the circular structures. The shape of a molecule can therefore be roughly predicted based on the number of valence electrons available in each atom.

    . .   . .
HÅ   ÅFl: = H—Fl :

Conformations

Structural formulas printed on a page make no attempt to illustrate the geometry or movement of a molecule. Molecules are moving, three dimensional structures and it is important to consider the rotational capabilities of the groups which make up a molecule. Groups bonded by single bonds are capable of spinning around the bond, much like a wheel spins around its axle. The different orientations in which a molecule may appear at any given time are its conformations. Conformations are not steady—one is always changing into another—and only give a view of a molecule as a “snapshot” of a moment in time. Conformations are different from isomers because the bond types and sequences do not change, only the orientation of the group at the end of the single bond.

Schrödinger equation

Austrian physicist Erwin Schrödinger proposed an equation describing the time dependence of quantum mechanical systems. It is a central factor in the theory of quantum mechanics and has been compared the Newton's second law in classical mechanics. Each system in the mathematical formulation of quantum mechanics is in concert with a complex Hilbert space, which is a generalization of Euclidean space and not restricted to finite dimensions, such that every instantaneous state of the system is described in that space by a unit vector. The vector encodes outcomes of all the possible measurements that are applied to the system. The vector becomes a function of time as the state of a system changes over a period of time. The equation is:

$$i\hbar\frac{\partial\Psi(\mathbf{r},\,t)}{\partial t} = \hat{H}\Psi(\mathbf{r},\,t)$$

In this, i is the unit imaginary number; ħ is Planck's constant divided by 2π, and the Hamiltonian H(t) is a self-adjoint operator that acts of the state space.

Aufbau principle

The Aufbau principle is a set of rules that help solve some of the modern atomic theory equations. Quantum numbers play a part in these rules. With each principal number n is an energy level called a shell. A shell 1 is n = 1, shell 2 is n = 2, etc. Each shell has subshells and can have more than one subshell. The number of subshells within a shell = n. Electrons reside in orbitals. The Aufbau principle says that the physical and chemical properties of elements are determined by the atomic structure. The principle rules for placing electrons within shells is known as the Aufbau principle. The rules are:

- The least energetically available subshell is where electrons are placed.
- At most, only two electrons can be held in an orbital.
- If two or more equal orbitals are available then these electrons should be dispersed before being paired.

Pauli exclusion principle

Wolfgang Pauli formulated the quantum mechanical principle which states that no two electrons within an atom may have quantum numbers that are identical. This also applies to fermions, which have half-integer spins. The building blocks of matter -- electrons, protons and neutrons -- are all subject to the Pauli exclusion principle. Many of the characteristic properties of matter, ranging from large-scale stability of matter to the existence of the periodic table of elements, are subject to the Pauli exclusion principle. These fermions have angular momentum and are considered antisymmetric states within quantum mechanics. Bosons are particles that are not fermions and have particles that use symmetric states in quantum theory. Bosons may only share quantum states and have integer spins such as the photon.

Planck's Constant

Planck's Constant can be shown by taking a rotating, charged conducting ball and slicing it perpendicular to its axis of rotation into tiny cylinders. Each of the cylinders will behave as a current loop. What this means is that a rotating charged ball has a moment when it is magnetic. The electrons and protons are not rotating in conducting charged balls but are modeled as spinning charges; they also become magnetic Their momentum is angular, called "spin," and quantized in units of h. Planck's Constant = 6.626 x 10-34 J x s. In a magnetic field the electron or proton will have spins of only h/2 or -h/2, which are up or down depending on direction of the field.

Rutherford scattering

Ernest Rutherford concluded from the work of Geiger and Marsden that the majority of the mass was concentrated in a minute, positively charged region, or the nucleus, which was surrounded by electrons. When a positive alpha particle approached close enough to the nucleus it was strongly repelled; enough so that it had the ability to rebound at high angles. The small nucleus size explained the small amount of alpha particles that were repelled in this fashion. Rutherford scattering is also referred to as Coulomb scattering. This is because it relies on static electric, or Coulomb forces. The scattering led to development of the orbital theory of the atom, or what is known as the Bohr model. Rutherford was known as the "father of nuclear physics."

Bragg's law

W.H. and W.L Bragg, father and son, derived from experiments of X-ray diffraction of crystal surfaces at certain angles that real particles at the atomic scale existed. The experiments also provided a new way in which to study crystals in the form of X-ray diffraction. When an atom is hit by an X-ray, they make the electronic clouds move (as does other electromagnetic waves). This movement helps re-radiate waves with the same frequency, which is known as elastic or Rayleigh scattering. These re-emitted X-rays limit constructive or destructive interferences. The interference is constructive when a shift in phase is proportional to 2π. This expresses Bragg's law:
In this, n is an integer; λ is the X-ray wavelength; d is the spacing between the planes in the atomic lattice; and θ is the angle between the scattering planes and incident rays.

Bohr model

The Bohr model is a model of the atom that gives an explanation about matters such as line spectra. Neil Bohr postulated that the electrons which orbit an atom could only occupy certain orbits. Those were orbits in which angular momentum satisfied a particular equation:

$$L_n = mv_n r_n = nh / 2\pi \qquad (n = 1, 2, 3, \ldots)$$

Where m is the electron's mass, r is the radius of the orbit and v the orbital speed of the electron. Bohr proposed that the angular momentum of an electron is quantized. To analyze this, take the concept of circular motion and the potential energy of two charges. The electron has a charge of -e, while the nucleus has a charge of +Ze, where Z is the atomic number of the element. The energy is then:
The electron has uniform circular motion

$$\text{Energy of an electron} : E = KE + PE = 1/2\, mv^2 - kZe^2 / r$$

with the only force on the electron being an attraction between the negative electron and the positive nucleus. So:
Rearrange the angular momentum equation to solve for velocity:
Substituted in the nth level:

De Broglie waves

Louis de Broglie was a French physicist who introduced electron wave theory. This included the wave-particle duality theory of matter which assumes that light and matter can show properties of both waves and particles. His work led to the hypothesis:

"Any moving particle or object has an associated wave."

This created wave mechanics physics, which joined light and matter physics. De Broglie helped develop an explanation of wave mechanics that was unlike other models involving probability, which are dominant in quantum mechanics. This is called the de Broglie-Bohm theory after being improved by David Bohm. The interpretation by Bohm holds that the existence of a non-local universal wave function allows distant particles to instantly interact. The wave function grows according to Schrödinger's equation and directs the particle.

Davisson-Germer experiment

The Davisson-Germer experiment exhibits the wave nature of the electron and confirmed de Broglie's earlier hypothesis. This helped give quantum mechanics a boost due to putting wave-particle duality on a stable experimental footing. Bragg's law for diffraction had been applied to the diffraction of X-rays, but this was the first application for particle waves. A vacuum apparatus was built by Davisson and Germer in order to measure the energies of electrons scattered from metal surfaces. Electrons from a heated filament were accelerated by voltage and allowed to strike the nickel metal's surface. The electron beam was aimed at the nickel which could be rotated to see angular dependence of the scattered electrons. The electron detector was mounted on an arc in order to rotate for observing electrons at different angles. They found a peak in the intensity of scattered electron beams at certain angles.

Copenhagen interpretation

The Copenhagen interpretation was formulated by Niels Bohr and Werner Heisenberg in Copenhagen circa 1927. It extended an earlier probabilistic interpretation that Max Born had proposed. The Copenhagen interpretation assumes that two processes influence the wave function. First is the unitary evolution based on the Schrödinger equation. The other is the process of the measurement. The first process can be interpreted that a wave function will have a wave function collapse in the second stage, or it can be imagined that the wave function serves as an alternate mathematical tool rather than a physical entity whose physical meaning is the ability for one to calculate the probabilities. Bohr said that only results of experiments should be predicted, thus more questions are philosophical rather than scientific.

Many-worlds interpretation

The many-worlds interpretation in quantum mechanics theory reject a non-deterministic and wave function collapse associated with the measurement in the Copenhagen interpretation. It favors a description along the lines of quantum entanglement -- where the quantum states of two or more objects are described as: referred to each other even though the objects are spatially separated -- and a reversible time of evolution of states. This is associated with measurement explained by the process of decoherence, in which a system interacts with its environment in a way that the different portions of its wave function may no longer interfere with each other. The many-worlds interpretations postulate that a state function for the universe obeying Schrödinger equation has no process of non-deterministic wave function collapse. Also theorized is that the universal state is a quantum superposition of perhaps infinite parallel universes.

Photoelectric effect

The photoelectric effect is a process in which light falling on a surface removes electrons from the surface. Einstein explained this effect and it was one of the earliest applications of quantum mechanics. To understand the effect, consider that light behaves like photons. Each electron is ejected by only one photon striking the surface. Quantum theory holds that frequency, f , of the light determines energy, E, of the photon in the light beam:

$$E = hf$$

where h is Planck's constant. The energy from the emitted electron is given by the photon's energy minus the energy that is required to release the electron from the surface. So, it depends on the light frequency that falls on the surface but not its intensity.

Double-slit experiment

The double-slit experiment lets light diffract through two slits and produce fringes on a screen. These fringes, or interference patterns, have light and dark regions that correspond to where light waves interfered constructively and destructively. The experiment may also be performed with a beam of atoms and electrons that show like interference patterns. This supports wave-particle duality predicted by quantum mechanics. But, a double-slit experiment may also be accomplished with water waves in a ripple tank. The explanation of the wave seen does not require quantum mechanics. Rather, the phenomenon is quantum mechanical only when quantum particles such as electrons, photons or electrons are manifested as waves.

Wheeler's delayed choice experiment

In 1983 John Archibald Wheeler proposed an experiment that was a variation of the double-slit experiment in which the target detector can be changed at the last moment through a delayed choice of the observer. There is a detector screen which detects the normal wave interference pattern producing the double slit which may be removed. Two tightly-focused telescopes are behind the screen and are pointed to see one slit or the other, detecting the path the photon traversed. If photons are observed with the screen, a wave-like behavior happens. If the photons are seen with the telescopes then particle-like behavior takes place. But there is another scenario, when the photon has originated from the other side of the universe and is bent by an intervening black hole so that it arrives by one or two different paths to the detector.

Quantum eraser experiment

The quantum eraser experiment is a double-slit experiment in which selective polarization and particle entanglement is used to determine into which slit a particle will go by measuring the particle's entangled partner. The partner never enters the experiment. The measurement of direction negates the interference pattern in the double-slit portion of the experiment. One can restore the interference pattern without changing anything in the double-slit experiment by destroying the directional information in the entangled particle that previously determined the path. This makes the experiment advantageous. The quantum eraser erases the directional information and restores interference without changing the double-slit experiment. The experiment can be modified, whether to measure or destroy the directional information; or can be delayed until after the entangled particle partner has/has not interfered with itself.

Quantum indeterminacy

Quantum indeterminacy is an apparent needed incompleteness in describing a physical system that is one of the characteristics of the quantum physics description. Before quantum physics, it was believed that a physical system with a determinate state specifically determines all the values of its properties that are measurable. Also prior to quantum physics, the values of its measurable properties were thought to uniquely determine the state. Quantum indeterminacy may be characterized by a probability distribution on the set of outcomes of measurements of an observable. The distribution is specifically determined by a system's state and quantum mechanics provides a way to calculate this probability distribution. This indeterminacy can be characterized quantitatively by a probability distribution on the set of outcomes of measurement of an observable, or a property that can be determined through a sequence of physical operations.

Tetraneutrons

A tetraneutron is a hypothetical cluster of four neutrons. This particle cluster is not supported by the present standard model of particle physics. But some empirical evidence suggesting this particle's existence exists based on an experiment by Francisco-Miguel Marques. The experiment used a method of watching the disintegration of beryllium and lithium nuclei using a particle accelerator to fire the atomic nuclei and produce a spray of particles from the collision. This is a potentially important discovery since nuclear theory suggests that the clusters should not be stable and therefore should not exist. Recent modifications have been made in experiments and if a bound tetraneutron can be confirmed it could significantly change the understanding of nuclear forces.

Weak interaction nuclear force

The weak interaction nuclear force is one of four fundamental forces of nature and is most commonly seen in beta decay and radioactivity. The word weak comes from the fact that the field strength is some 1013 times less strong than a nuclear force. This weak force affects all leptons and quarks and is the only force that affects neutrinos, a fermion. This weak interaction lets all quark and lepton particles and antiparticles interchange mass, energy, electric charge, and flavor (a quantum number of elementary particles related to their weak interactions effectively changing into one another). Their large mass is approximately 90 GeV/c2 and their mean life is only about 3 x 10-25 seconds by the uncertainty principle.

Standard Model of particle physics

The Standard Model is a theory developed in the 1970s describing strong, weak and electromagnetic forces as well fundamental particles that comprise matter. It is consistent with both special relativity and quantum mechanics. It is not a complete theory of fundamental interactions mainly because it does not give a description of gravitational force. The model contains both fermionic and bosonic particles. The fermions are basically particles of matter and bosons are transmission force particles. The theory holds that photons mediate electromagnetic interaction, W and Z bosons mediate weak nuclear force, eight gluon species mediate strong nuclear force, and Higgs bosons cause spontaneous symmetrical breaking of the electromagnetic and weak interaction gauge groups and are a reason for inertial mass to exist.

Spallation

Spallation is the process in which heavy nuclei emit large numbers of nucleons due to being struck by high-energy protons, greatly reducing its atomic weight. Spallation takes place naturally in the Earth's atmosphere due to impacts of cosmic rays as well as on the surfaces of bodies in space, such as the moon and meteorites. Cosmic ray spallation is evidence that materials have been exposed on the surface of the body and, partly, gives a way to measure the length of time exposed. The composition of cosmic rays also indicate that they have undergone spallation before reaching the Earth (due to the proportion of such light elements as Li, B and Be which exceed the average cosmic abundances). The elements in cosmic rays were probably formed from oxygen, nitrogen, carbon, and maybe silicon from cosmic ray sources or during lengthy travel to Earth.

Electromagnetic radiation

Radiation is heat transferred by electromagnetic waves that are emitted and carry energy away from the object that omits them. Two bodies of heat do not need a medium for transferring heat in radiation such as is needed by convection and conduction. Instead, photons traveling at the speed of light act as the intermediaries. Heat transferred into or out of something has components that include emissivity, surface area, surface reflectivity temperature and geometric orientation. An object's emissivity and surface reflectivity are functions of both the composition and surface condition. Radiation as a means of transferring heat must account for thermal heat that is both incoming and outgoing.

Electromagnetic radiation is energy in the form of electromagnetic waves which consist of both a magnetic field component and an electrical field component. Electromagnetic waves travel at a velocity of 3.0 × 108 meters per second, the speed of light in a vacuum. There are various types of electromagnetic radiation which can differ from each other in wavelength and frequency. The wavelength and frequency are related by:

$$\nu = \frac{c}{\lambda}$$

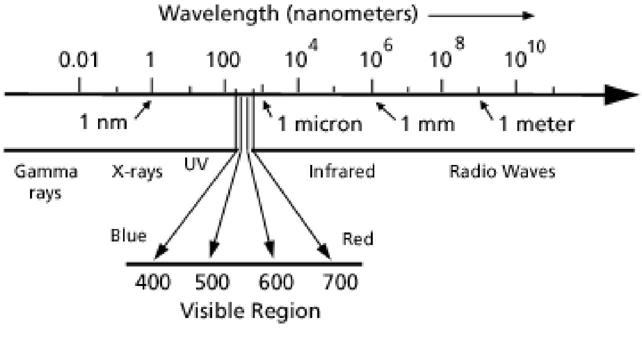
where c = speed of light; λ = wavelength; v = frequency.

Electromagnetic spectrum

The electromagnetic spectrum is the range of all possible radiation that is electromagnetic. The spectrum is also the range of electromagnetic radiation that is reflected, transmitted or emitted. The spectrum extends from the long-wavelength end, for frequencies used in power grids, to gamma radiation at the short wavelength end. It covers wavelengths from thousands of kilometers down to sizes that are fractional of atoms. Sunspots produce radiation every 22 years or a frequency of 1.4 x 10-9 Hz. Photons, at the other end, with arbitrarily high frequencies may be made by electrons colliding with positrons at appropriate energy. Photons of 1024 Hz can now be produced by man-made accelerators.

Types

Electromagnetic radiation differs in energy level and each type is spread over a range of wavelengths. The lowest energy level radiation is radio waves. These have the longest wavelengths (~1010 nm) and lowest frequency. The highest are gamma rays which have very short wavelengths (~ 10-3 nm) and high frequencies. The wavelength of the visible light region ranges from ~ 400 nm (violet) to ~ 700 nm (red). Between these two are the infrared spectrum, the UV spectrum, visible light and X-rays.



Electromagnetic Spectrum

Cherenkov radiation

Cherenkov radiation is electromagnetic radiation that is emitted when charged particles pass through insulators at speeds greater than that of light in the particular medium. The so-called "blue glow" that is known in nuclear reactors stems from Cherenkov radiation. It occurs because relativity holds that the speed of light in a vacuum is a universal constant (c), and the speed of light in a material may be much less than c. The speed of light in water is just 0.75c, for example. Cherenkov radiation occurs when charged particles, usually electrons, exceed the speed of light in an electrically insulated medium through which it travels. The intensity of Cherenkov radiation is in proportion to the velocity of the inciting particle and to the numbers of those particles. This radiation has higher frequencies and thus shorter wavelengths. Cherenkov radiation is used in detecting high-energy charged particles.

Thermal radiation

Thermal radiation is electromagnetic radiation from objects that are caused by its temperature. It has rapid increases in power and frequency along with increased temperature. Radiant heat is used in homes, such as when electricity is forced through materials for conductive heating. Black bodies are objects that take in all electromagnetic radiation falling on it. Heat radiators are used, such as those mounted on the interior surface of payload bay doors in the Space Shuttle. The shuttle can orbit and still allow the radiators to work. These radiators are used as black body radiators, but the coatings appear white to the eye because of the need to reflect wavelengths that are visible in order to avoid their absorption.

Electromagnetic wave equation

The electromagnetic wave equation is a second-order partial differential equation that guides the propagation of electromagnetic waves through a medium or in a vacuum. The equation written in terms of either the electric field, E, or the magnetic field, H, is such that:

$$\nabla^2\mathbf{E} - \frac{1}{c^2}\frac{\partial^2\mathbf{E}}{\partial t^2} = 0$$

$$\nabla^2\mathbf{H} - \frac{1}{c^2}\frac{\partial^2\mathbf{H}}{\partial t^2} = 0$$

where c is the speed of light in the medium. In a vacuum, c = 2.998 x 108 meters per second which is the speed of light in free space. The electromagnetic wave equation is derived from Maxwell's equations. In a linear isotropic, not-dispersive medium, the magnetic flux density, B, relates to the magnetic field, H, by:

B = oμH

where ημ is the magnetic permeability of the medium; B is normally called the magnetic field in most modern literature and H is the "H vector" or auxiliary magnetic field.

Hazards

Electromagnetic radiation can be either ionizing radiation or non-ionizing radiation depending on whether it is able to ionize atoms and break chemical bonds. Frequencies such as ultraviolet and higher, such as gamma rays or X-rays, are ionizing. These have their own potential dangers. Non-ionizing radiation is not capable of having such effects on molecules, at least in this context. Three potential hazards that are considered major exist with non-ionizing electromagnetic radiation. They are electrical, biological and fire. The induced current caused by radiation also can pose significant dangers in handling pyrotechnics. Induced current can cause electric shock to humans or animals. A strong electromagnetic field may cause electric currents which flow over air gaps to the ground, causing sparks and fire. The biological hazard stems from electromagnetic fields causing dielectric heating and can lead to tissue and other bodily damage.

Infrared radiation

Infrared radiation is a form of electromagnetic radiation with a wavelength longer than visible light but shorter than microwave radiation. The name stems from the Latin "infra," meaning below and red because it is the color of visible light of longest wavelength. Three orders of magnitude are spanned by infrared radiation and it has wavelengths that are from 750 nm to 1 mm. Infrared has divided regions: near infrared, defined by water absorption and used commonly in fiber optic telecommunications because of low losses in the glass medium; short wavelength or short wave, in which water absorption significantly increases at 1,450 nm; mid wavelength; long wavelength and far infrared.

Night-vision equipment: Night-vision equipment such as that used by military, police and firefighters uses infrared radiation when not enough visible light is available to see an object. Radiation is detected and turned into an image on a screen. Objects with higher temperatures show up in different shades than cooler objects which allows police and military to find important thermal targets such as humans and battle tanks. Infrared radiation is not heat itself but rather another effect of heat. Thermal detectors do not find heat directly but instead find the difference in the infrared radiation of objects. Certain materials emit more or less infrared radiation upon increase or decrease of temperature, although it depends on the materials' composition. Firefighters use infrared imaging in smoke-filled spaces because smoke is more transparent to infrared than it is to visible light.

Radioactive decay

Radioactive decay is a set of processes that allow unstable atomic nuclei, or nuclides, to emit subatomic particles, or radiation. The decay is a random process and it is not possible to predict an individual atom's decay. The becquerel, Bq, is the SI unit for measuring radioactive decay. One Bq is defined as one decay event per second. Certain interactions of protons including strong nuclear force, weak nuclear force and electrostatic force play a part in decay. Some particle configurations in a nucleus have properties that if they were to shift slightly those particles could be considered a lower-energy structure. A decay event needs activation energy. Quantum mechanical particles continuously and randomly move. So, if constituent particles move together, the nucleus can destabilize spontaneously.

Beta decay

Beta decay is radioactive decay in which there is an emission of a beta particle, a high-energy electron or positron. In electron emission it is known as beta minus (β-). In positron emission it is known as beta plus (β+). In β- decay there is conversion of a neutron into a proton caused by the interaction of the electron is emitted with an anti-neutrino. This is because of the conversion of a down quark (first-generation quark with a charge of 1/3e) to an up quark (a first-generation quark with a charge of +(2/3)e by a W boson emission. In β+ decay, a proton is converted into a neutron, a positron and a neutrino:

$$\text{energy} + p^{+} \rightarrow n^{0} + e^{+} + \nu_{e}$$

Double beta decay

In double beta decay, unstable nuclei decay by the conversion of a neutron in the nucleus to a proton and emitting an electron and anti-neutrino. Beta decay must have a greater binding energy than the original nucleus. With some nuclei such as germanium-76, the nuclei, with an atomic number of one higher, has a smaller binding energy that prevents beta decay from happening. With germanium-76, the nuclei with an atomic number two higher, selenium-76, has greater binding energy so the process of "double beta" decay can take place. In double beta decay, two neutrons in the nuclei convert to protons and two electrons and two anti-neutrinos are omitted. It is the rarest type of radioactive decay that is known.

Gamma rays

A gamma ray (γ)is electromagnetic radiation that is produced from radioactive decay or other nuclear processes such as an electron annihilation, in which an electron and positron collide. No physical difference exists between X-rays and gamma rays that have the same energy, they are set apart by their origin. Gamma ray describes high-energy electromagnetic radiation by nuclear transitions while X-rays are high-energy radiation caused by energy transitions from electrons that are accelerating. Gamma rays penetrate more than either alpha or beta radiation, both of which are not electromagnetic radiation, but is not as ionizing. Gamma sources have a wide variety of uses ranging from medicine to industry.

Spontaneous fission

Spontaneous fission, or SF, is a form of radioactive decay that is typical of very heavy isotopes. This fission is possible, in theory, for any atomic nucleus with a mass greater than or equal to 100 amu. In reality, however, spontaneous fission is only energetically possible for atomic masses above 230 amu. Most susceptible to spontaneous fission are those trans-actinide elements such as rutherfordium. Spontaneous decay by fission does happen for uranium and thorium but it is not observed for the majority of radioactive breakdowns. The mathematical formula for spontaneous fission occurring is:

$$Z^2/A \approx 45.$$

Spontaneous fission goes through the same process as nuclear fission only it does not generate a neutron flux necessary for going "critical" and continuing such functions. The other distinction between spontaneous and nuclear fission is the former is not self-sustaining.

Isomeric transition

Isomeric transition, a radioactive decay process, happens in an atom when its nucleus is in an excited meta-state, such as after the emission of alpha or beta particles. The nucleus has extra energy released by gamma ray emission that returns the nucleus to the ground state. Gamma emission is similar, but it involves excited meta-states. After decay, the nucleus remains in an excited state. When energy is released as gamma rays there is no change in Z or A = isomeric. Gamma rays may transfer energy directly to tightly-bound electrons that cause ejection from an atom in a process called internal conversion.

Internal conversion

Internal conversion is a radioactive decay process where gamma rays emitted from a nucleus are photoelectrically absorbed by one of the most tightly bound electrons. This causes it to be ejected from the atom. Following ejection of the internal conversion electron, another shell electron fills the vacancy along with a corresponding emission of one or more X-rays or an auger electron, the latter of which produced by an auger emission in which an electron of an atom causes the emission of a second electron; this second electron ejected is the Auger electron. Favorable circumstances for internal conversion are when energy gaps between nuclear levels are small; this is the only way for de-excitations of 0+->0+ transitions can occur.

Electron capture

Electron capture, a decay mode for isotopes, occurs when too many protons are in the nucleus of an atom and there is not sufficient energy to emit a positron. This will continue to be an able decay mode for radioactive isotopes that can decay by positron emission. If the difference in energy between the parent and daughter atom is less than 1.022 MeV, positron emission cannot happen and electron capture is the only mode of decay. One example is Rubidium-83 decaying to Krypton-83, which can only happen by electron capture. In such a case, one of the electron's orbitals (normally from the K or L electron shells) is captured by a proton in the nucleus and forms a neutron and a neutrino.

Nuclear reaction

A nuclear reaction is a process in which two nuclear particles or nuclei collide, producing different products than the original. A reaction can encompass more than two colliding particles but such events are very rare. If particles collide and separate without undergoing change, the process is termed a collision instead of a reaction. It can be written with an equation in which each particle in the reaction is written with its chemical symbol then atomic number subscripted as well as a superscripted atomic mass. Neutron is n, electron e, and the proton may be represented by H or p. The sum of the atomic numbers must be on each side of the equation in order to be balanced; the sum of the atomic masses must also be on each side and are also equal. Such as:

$$^{6}3\text{Li} + {}^{2}1\text{H} \rightarrow {}^{4}2\text{He} + ?$$

The balanced equation for the helium nucleus is:

$$^{6}3\text{Li} + {}^{2}1\text{H} \rightarrow {}^{4}2\text{He}$$

Release of energy

Energy in a reaction can be calculated according to a reference table of very accurate particle masses. For instance, the reference tables indicate that the 63Li nucleus has an atomic weight of 6.015 amu, the deuteron is 2.014 amu and the nucleus is 4.0026 amu. So, the left side total mass is 6.015 + 2.014 = 8.029; right side total mass is 2 x 4.0026 = 8.0052; the missing mass is 8.029 - 8.0052 = 0.0238 amu. The mass that is missing stems from energy that is released from the reaction. It's source is the binding nuclear energy. To determine how much energy has been released, Einstein's E=mc2 can be used. One amu = 931 MeV, so energy released is 0.0238 x 931 MeV = 22.4 MeV.

Neutron radiation

Neutron radiation consists of free neutrons. The neutrons may come during nuclear fission, nuclear fusion or from very high energy reactions as well as other sources. Neutron radiation is often called indirectly ionizing radiation. It ionizes atoms differently than neutrons having no charge. But interactions are largely ionizing such as when neutron absorption causes gamma emission and the gamma rays remove electrons from an atom or a nucleus recoils from neutron interaction and is ionized, causing more traditional ionization in other atoms. Neutrons are more penetrating than alpha or beta radiation because they are not charged and they sometimes penetrate more than gamma radiation. Neutron radiation has the ability to induce radioactivity in most substances it touches including body tissues of humans. This happens through neutrons captured by atomic nuclei and transformed in radionuclide. This process is the basis for much of the radioactive material from detonated nuclear weapons.

Nuclear meltdown

A nuclear meltdown happens whenever nuclear reactor cores melt. It usually is thought of as a serious nuclear accident. The reactor core may melt through the reactor floor chamber and drop down, continuing downward until it melts so that surrounding material dilutes and cools to a temperature that is no longer hot enough to cause melting through the material underneath, or until it hits groundwater. A thermonuclear explosion does not happen in meltdowns, but a steam explosion can happen if water is hit. Geometry and water is important in maintaining the chain reaction so the molten core cannot form uncontrolled critical mass. Molten reactor cores continue generating heat that is sufficient through radioactive decay so that it maintains or even increases water temperature.

Light water nuclear reactors

Nuclear fission reactors that are used for electric power generation are considered "light water reactors." Light water, which is ordinary water, is used as the moderator in American reactors. The water is also used as a cooling agent and the heat is taken away to make steam for moving turbines in electric generators. Ordinary water use requires a certain amount of uranium fuel enrichment before the criticality of the reactor is maintained. Two types of light water reactors exist. Pressurized water reactors are those in which water passes over the reactor core as a moderator and coolant does not flow to the turbine but is instead kept in a pressurized loop. Water passes over the reactor core in a boiling water reactor in order to act as moderator and coolant also acts as the source of steam for the turbine.

Fast breeder reactors

Neutrons given off by fission reaction can, under certain circumstances, breed more fuel from isotopes than would otherwise be non-fissionable. Plutonium-239 is the most common breeding reaction from non-fissionable uranium-238. Uranium-238 is bombarded with neutrons that set off two successive beta decays with plutonium production. How much plutonium is made is dependent on the breeding ratio. Fast breeder relates to the configuration types which actually produce more fissionable fuel than they use. An example is a liquid metal fast-breeder reactor. The reason for this is because non-fissionable uranium-238 is some 140 time more abundant than fissionable U-235 and can be converted efficiently into Pu-239 by neutrons from a fission chain reaction.

Periodicity

Periodic table

The periodic table is a tabular arrangement of the elements and is organized according to periodic law. The properties of the elements depend on their atomic structure and vary with atomic number. In the periodic table, the elements are arranged by atomic number in horizontal rows called periods and vertical columns called groups or families. They are further categorized as metals, metalloids, or nonmetals. The majority of known elements are metals; there are seventeen nonmetals and eight metalloids. Metals are situated at the left end of the periodic table, nonmetals to the right and metalloids between the two.

Periodicity

Periodicity describes the predictable and incremental nature of elements' properties and places them on the periodic table accordingly. An atom of every element has unique properties such as number of electrons, density, and mass. The periodic table is arranged such that elements near each other are more alike in these properties than those that are far apart on the table. Periodicity enables the prediction of properties and atomic configurations based on known trends represented by the position of elements on the table. One such trend is the number of electrons; reading from left to right on any given row of the table, each element contains one more electron than the one immediately preceding it. On row 4, for example, K contains 19 electrons and the next element, Ca, contains 20.

Important information and how to read

Most periodic tables contain the element's atomic weight, number, and symbol in each box. The position of an element in the table reveals its group, its block, and whether it is a representative, transition, or inner transition element. Its position also shows the element as a metal, nonmetal, or metalloid. For the representative elements, the last digit of the group number reveals the number of outer-level electrons. Roman numerals for the A groups also reveal the number of outer level electrons within the group. The position of the element in the table reveals its electronic configuration and how it differs in atomic size from neighbors in its period or group. In this example, Boron has an atomic number of 5 and an atomic weight of 10.811. It is found in group 13, in which all atoms of the group have 3 valence electrons; the group's Roman numeral representation is IIIA.

Features and structure

The most important feature of the table is its arrangement according to periodicity, or the predictable trends observable in atoms. The arrangement enables classification, organization, and prediction of important elemental properties. The table is organized in horizontal rows called Periods, and vertical columns called Groups or Families. Groups of elements share predictable characteristics, the most important of which is that their outer energy levels have the same configuration of electrons. For example, the highest group is group 18, the noble gases. Each element in this group has a full complement of electrons in its outer level, making the reactivity low. Elements in periods also share some common properties, but most classifications rely more heavily on groups. A typical periodic table shows the elements' symbols and atomic number, the number of protons in the atomic nucleus. Some more detailed tables also list atomic mass, electronegativity, and other data.

Elements and blocks

Elements are arranged in blocks on the table in addition to being classified by periods and groups. The two groups at the left of the table are the s block, groups 3-12 are the d block, groups 13-18 are the p block, and periods 4f and 5f are the f block. Representative elements are those occupying the s and p blocks. Transition elements are those occupying the d block. Inner transition elements are those occupying the f block. The names, "representative", "transition", and "inner transition" refer to the electron configuration of each element's atoms. The representative elements each have from one to six electrons in the outer energy levels. The transition elements have electron configurations that describe the filling of different sublevels, thereby transitioning from the s block to the p block.

Chemical reactivity

Reactivity refers to the tendency of a substance to engage in chemical reactions. If that tendency is high, the substance is said to be highly reactive, or to have high reactivity. Because the basis of a chemical reaction is the transfer of electrons, reactivity depends upon the presence of uncommitted electrons which are available for transfer. Periodicity allows us to predict an element's reactivity based on its position on the periodic table. High numbered groups on the right side of the table have a fuller complement of electrons in their outer levels, making them less likely to react. Noble gases, on the far right of the table, each have eight electrons in the outer level, with the exception of He, which has two. Because atoms tend to lose or gain electrons to reach an ideal of eight in the outer level, these elements have very low reactivity.

Groups and periods

Reading left to right within a period, each element contains one more electron than the one preceding it. (Note that H and He are in the same period, though nothing is between them and they are in different groups.) As electrons are added, their attraction to the nucleus increases, meaning that as we read to the right in a period, each atom's electrons are more densely compacted, more strongly bound to the nucleus, and less likely to be pulled away in reactions. As we read down a group, each successive atom's outer electrons are less tightly bound to the nucleus, thus increasing their reactivity, because the principal energy levels are increasingly full as we move downward within the group. Principal energy levels shield the outer energy levels from nuclear attraction, allowing the valence electrons to react. For this reason, noble gases farther down the group can react under certain circumstances.

Atomic radius

The radius of an atom is basically understood as the distance between the center of its nucleus and the outer energy level of electrons. Atomic radii are of course incredibly small and are measured in nanometers (1nm = 10^{-9} m). The radius of an element tends to increase reading downward within a period because the outer energy electrons are further from the nucleus due to the shielding effect of lower energy levels. The radius of an element tends to decrease reading from left to right within a period because of the increasing charge. Recall that each subsequent element in a period has one additional proton and one additional electron. Since the positive charge of the nucleus and the negative charge of the outer energy level have both increased, with no increase in shielding, the results are a stronger attraction and a smaller radii.

Electron configuration and ionization energy

Electron configuration is a description of how an atom's electrons are distributed throughout its energy levels and sub-levels. The energy level, sub-level, and number of electrons are represented with numbers and letters. For example, H has one electron in one energy level; its electron configuration is represented: $1s^2$. More complex atoms with higher atomic numbers have more complex configurations and notations, but the configuration is predictable based on the element's location on the periodic table. Ionization energy, or potential, is the energy required to separate electrons from their atoms, creating ions. Electrons will not separate from the atoms to which they are bound without some force acting upon them. Ionization energy is one of the trends exhibited in the periodic table; it decreases reading from top to bottom within a group, and increases reading from left to right within a period.

Electronegativity

The electron attracting power of an atom is its electronegativity and this determines a large part of its chemical behavior. Electronegativity is measured on a scale from 0 to 4 (weakest to strongest). The most electronegative elements are listed to the right of the periodic table. Conversely, those elements with low electronegativities are found at the left of the periodic table. In general, electronegativity decreases going down each group in the periodic table.

Electronegativity refers to the ability of an atomic nucleus to attract electrons when covalently bonded in molecules. In certain circumstances, one atom of a molecule may exert a stronger attraction upon a shared pair of electrons than the other atom sharing the same pair. The unequal sharing results in the shared pair being pulled closer to the nucleus of the "stronger" atom. For example, in a molecule of hydrogen chloride, the shared electrons are pulled closer to the nucleus of the Cl atom, resulting in a slight negative charge of the atom and a corresponding slightly positive charge of the H. The shared electrons do not leave the covalent bond; they are merely pulled closer to the Cl nucleus showing greater electronegativity in Cl than in H. Electronegativity is a trend predicted by the periodic table; it increases reading upward within a group and left to right within a period.

The Pauling scale is a system of measuring electronegativity (written by Linus Pauling in 1932). The scale assigns ascending numerical values to represent electronegativity, with the lowest at .7 (francium) and the highest at 3.98 (fluorine). Simple subtraction yields the difference between electronegativity of elements in a compound (represented as σEN), which in turn allows predictions about types and strengths of bonds likely to be formed. Ionic bonds are characterized by large σEN values, while covalent bonds will have very little to no difference in electronegativity. The Mulliken scale is another system for measuring electronegativity (devised by Robert Mulliken in 1934). Rather than arbitrarily assigning beginning and ending points, however, this scale's electronegativity values are the average of an element's ionization potential and electron affinity. The result is shown in electron volts.

Chemical bonds

Most chemical bonds can be classified as either ionic or covalent. Metals typically contain fewer valence electrons than nonmetals and form ionic bonds with other metals. The type of bond likely to be present in a compound can be predicted by the position of the elements on the periodic table. The metals are found in groups 1-12 of the periodic table (though H is not a metal). Elements within these groups will tend to form ionic bonds with one another because fewer valence electrons are available for sharing. The nonmetals are found in the right 1/3 of the table, or groups 13-18. Because they have more valence electrons available for sharing, they tend to form covalent bonds both in bonding with other nonmetals and with metals.

Neutrons in atom models

Atoms that have similar outer electrons also have similar properties and their chemical properties take place on a repeating basis. The periodic table arranges the elements so that atoms with the same number of outer electrons are above each other in the table. The first column of the table, the metals, includes all the elements whose atoms have only one loosely bound outer electron. The last column, the noble gases, has all the elements whose atoms have no outer electrons. Protons and neutrons attract each other via a strong nuclear force. This force works only between those nucleons that are neighbors while an electric repulsion is long range but only works between protons. Adding neutrons to the nucleus increase the amount of strong force without an increase in the electric repulsion.

Elemental symbols

Many element symbols are the first letter of the element capitalized, such as “O” for oxygen, “S” for sulfur and “H” for hydrogen. Since more elements exist than letters of the alphabet and the names of several elements begin with the same letter, the symbolizing must be varied; The symbols may also be the first letter of the element capitalized followed by the second letter of the element such as “Al” for aluminum and “Ba” for barium. Since some of the elements have the same first and second letters, for example calcium and cadmium, the element symbol may be the first letter capitalized followed by the third letter of the element “Ca” for calcium or “Cd” for cadmium. Non-English words are also chosen as elemental symbols such as “K” for potassium, from the German “kalium”; “Fe” for iron, from the Latin “ferrum”; “Cu” for copper, from the Latin “cuprum”; and “Ag” for silver, from the Latin word “argentum.”

Bonding

Adhesion

The molecular attraction that is exerted between bodies when in contact is called adhesion. Five types of adhesion exist. Mechanical adhesion is when two materials become mechanically interlocked. An example is sewing. Chemical adhesion is when two materials form a compound at the joint. The strongest joints are where atoms from the two materials have an ionic bonding or covalent bonding of outer electrons. Dispersive adhesion, also known as adsorption, is when two materials are held together by what is known as "van der Waals forces." These forces are attractions between two molecules that have ends that are both positively and negatively charged. Electrostatic adhesion occurs when conducting materials pass electrons to each other forming a difference in electrical charge at the joint. The final type of adhesion his when materials merge at the joint by diffusion.

Chemical bonds

Chemical energy and bond energy

Chemical energy is the latent, or potential energy contained in atoms and molecules resulting from the forces within them. While the precise chemical energy of a given particle or substance cannot be determined, the difference in energy between one chemical state and another can be readily calculated and expressed as enthalpy. When chemical energy is lower, the particle is said to be in an unexcited, or stable state, and is less likely to create or sever bonds. Bond energy is the potential energy stored in chemical bonds. Chemical reactions involve the exchange of energy as bonds are created and/or broken. This exchange is measured in calories (an English system measurement) and joules (a metric unit of measurement). The heat and light produced by combustion of natural gas is an example of the release of chemical and bond energy.

Types and properties

Chemical bonds are the forces which hold atoms together in molecules and compounds. Three common types are:

- Ionic bonds: atoms are attracted to one another by electrostatic forces due to the opposite charges of their ions. Inorganic binary compounds composed of a metal and a non-metal are typically formed with ionic bonds because the metal loses an electron, gaining a positive charge, while the non-metal gains the electron, resulting in a negative charge.
- Covalent bond: atoms share, or swap, one or more pairs of electrons in their outer shells, producing a bond of roughly the same strength as ionic bonds.
- Metallic bond: a relatively weak bond formed by electrostatic attraction between a metal atom's core and its outer-level electrons. The outer energy levels of individual atoms essentially overlap, allowing valence

electrons to easily transfer from one atom to another. These weak bonds account for the low melting point and malleability of most metals.

Chemical reactivity

Chemical reactivity refers to the tendency of a given element or compound to be involved in chemical reactions. As chemical reactions are essentially the transfer of electrons to form or break chemical bonds, producing new elements or compounds, the bonds present between molecules are important factors in determining reactivity. Weak bonds and unstable (unbalanced) molecules lend themselves more readily to chemical reactions, while strong bonds and stable molecules are less likely to react without significant external activation energy. For example, hydrocarbons vary in the number of bonds between atoms of H and O; molecules in which the H and O are connected with double or triple bonds are less likely to react at that site. Other components of the molecule may be very reactive, but the double and triple bonds are stable and require a great deal of energy to break.

Molecular geometry

Molecules are groups of atoms bound by overlapping energy levels which have a specific orientation in three-dimensional space. The resulting molecule has a three dimensional shape which we can imagine as straight lines between the nuclei of bonded atoms. Molecular geometry, or molecular shape, refers to the characteristic shapes of molecules. These shapes are determined by the number and type of atoms present and the type(s) of bonds connecting them. A water molecule, for example, is bent, or v-shaped, with the two H atoms bonded at angles to a central O atom. The positions and angles of a molecule's bonds are predictable because electron pairs repel one another and arrange themselves around the central atom to be as far away from each other as possible. Common molecular shapes include tetrahedral, linear, angular, and pyramidal.

States of matter

All matter is known to have various forms, or states, of existence—solid, liquid, and gas. These states are in large part determined by the atomic properties of the elements and compounds and their bonding propensities. Atoms and molecules in a solid exist in a low energy state in which their movement is greatly decreased. They align themselves into consistent organizations in which intramolecular bonds play an important role. Because of the decreased kinetic energy, bonds in solids are relatively stable and require external energy to change them. Particles in a gaseous state move rapidly and randomly, increasing the chances that bonds will be made or broken. The kinetic energy of liquid particles is smaller than in gasses and greater than in solids. Because their motion is slower and less random, they condense into the characteristic form of liquids but do not align themselves so neatly as to form solids.

Ionization energy

The ionization energy of an atom or molecule, also called ionization potential, is the energy required to strip the atom or molecule of one electron. The nth ionization energy is the energy required to strip it of an nth electron, after the first, n - 1, is already removed. It is a measure in physical chemistry of the reluctance of an atom or molecule to give up an electron or the strength by which the electron is anchored. Generally, ionization energies decrease down a group of the periodic table and increase from left to right across a period. Ionization energy has a strong negative correlation with the atomic radius. Atomic ionization energy can be predicted by an analysis that uses the Bohr model and electrostatic potential, or the potential energy per unit of charge that is associated with a static electric field.

Chemical bonds

Molecular orbitals

Molecular orbitals describes the probable location of electrons in a molecule. Atomic orbitals are the probable locations of electrons within atoms. They are envisioned and represented as spherical, oval, and teardrop-shaped regions of space which together form the molecule’s geometry. They are formed by arranging constituent atomic orbitals in a linear plane with varying degrees of overlap to demonstrate sharing of electrons between atoms. The overlap of two atomic 1s orbitals, represented as spheres, produces an oval-shaped molecular orbital which is symmetrical around the two nuclei. The new orbital is called a sigma (denoted by the Greek letter σ) molecular orbital. When two 2p orbitals overlap, the resulting molecular orbital is not symmetrical, and is called a pi (π) molecular orbital. The larger the area of overlap, the stronger the resulting chemical bond. This theory partially explains the strength of CH bonds in hydrocarbons; they are σ molecular bonds.

Hybridization

Orbital hybridization is the mixture, or shifting, of ground state orbitals within an atom which permits the formation of the necessary number of bonds to meet the octet rule. This is a widely accepted theory developed by Linus Pauling; it has great explanatory power, but does not correspond to actual physical processes. The theory predicts that a 2s electron becomes energized, or excited, and is “promoted” to a vacant 2p orbital; then, the half filled 2s orbital and the three half filled 2s orbitals combine to form the hybrid sp3 orbital which can then form bonds with other atoms. The ability of C to form a virtually infinite number of compounds despite its ground state electron configuration of 1s22s22p2 is explained by the orbital hybridization theory. Though C is the most striking application of the theory, N and O can also hybridize.

Properties of C

Carbon is the most significant element in organic chemistry due to its abundance and bonding properties. C tends to form four covalent bonds, either with other C atoms or with other elements, because of its four valence electrons. This tendency is called tetravalency and accounts for the large number and strength of carbon compounds. These properties also explain the possible geometry of C compounds: if all four valence electrons form covalent bonds, the resulting molecule has a tetrahedral shape; this can be visualized as placing bonded atoms around the central C atom at the four corners of a three-dimensional pyramid. C is represented in structural formulas as C surrounded by four single lines on a plane; this structure is merely for convenience—one should always remember that molecules are three dimensional.

Bond dissociation energy

Simply stated, bond dissociation energy is the energy required to break a bond and is represented in equations as DH°. This energy can be measured and is useful in calculations involving enthalpy, the measurement of heat energy released or absorbed in a reaction. The bond in question is broken homolytically, meaning that each portion of the now-separated compound has one of the electrons that formed the covalent bond. The general process is represented as follows, where the double dots represent a pair of shared electrons: $X : Y \rightarrow X^{\cdot} + {}^{\cdot}Y$. In reality, however, the bonds under examination are parts of polyatomic molecules, so the bond dissociation energy is difficult to measure precisely. It increases as the difference in electronegativity between two atoms increases and is thus predictable using the trends on the periodic table.

Coordinate covalent bonds

If one atom in a binary compound or bonded pair provides both electrons for the covalent bond, a coordinate covalent bond results. For instance, when ammonia (NH3) bonds with boron trifluoride (BF3), the electron pair in the covalent bond between N and B come only from N. N has five valence electrons, three of which are bonded to H in the ammonia molecule. B has three valence electrons, all of which are bonded to F. The N—B covalent bond, therefore, depends on the two remaining electrons from N. Coordinate covalent bonds are important in a large number of biological processes involving essential metals such as zinc, magnesium, copper, and iron. They can have a large impact on molecular geometry because of the demanding nature of the conditions in which coordinate covalent bonds may be formed.

Bonding capacity of nonmetals

Nonmetals have predictable bonding capacities based on the number of valence electrons in an atom of each element. Because of their electron configurations, nonmetals form molecular compounds with other nonmetals through covalent bonds. Each nonmetal atom tends to form the number of covalent bonds required to meet the octet rule—a number corresponding to the number of single or unpaired electrons in its outer energy level. For example, C contains four valence electrons, so it tends to form four covalent bonds to meet the octet rule; O contains two valence electrons and forms two covalent bonds. The number of covalent bonds a nonmetal atom is likely to form can easily be determined from its dot structure. In the two examples shown here, the dot structure for fluorine shows one unpaired electron, so atomic fluorine will form one covalent bond. Selenium forms two.

..
: F :
.

.
: Se :
.

Bonding capacity of metals

Metals have very small numbers of valence electrons; most have two, and some have one or more than two. In a solid state, metals form strong bonds because their loosely bound electrons become delocalized, or they overlap with electron orbitals of neighboring atoms, much like a covalent bond. A solid metal is therefore a large collection of atoms whose electrons overlap with the orbitals of up to eight neighboring atoms, forming very strong bonds through the electrostatic attraction between positive nuclei and a sea of negative orbiting electrons. These strong bonds result in many of the typical metallic traits such as high melting point, ductile strength, and malleability. The delocalized electrons also account for the conductivity of metals—another characteristic property—because their weak attraction to a central atom means they are free to disperse electrical charges as they move about the solid.

Nonbonding electrons

Nonbonding electrons, also called lone-pair electrons or unshared electron pairs, are those in lower energy levels (non-valence) and those in valence shells not engaged in a bond at any given moment. Though not involved in bonds, they exert influence on the shapes of molecules by their forces of attraction and repulsion. Because they occupy more space than bonded electrons, nonbonded electrons exert greater force on neighboring atoms and molecules and can force the orbitals into closer contact than VSEPR theory would seem to predict, thus changing the geometry of the molecule. Atoms or molecules containing at least one unpaired electron are called free radicals; their lone electrons make them highly reactive and unstable. Nonbonding electrons in lower energy levels can become available for bonding through hybridization.

Diatomic elements

Diatomic elements tend to appear in pairs because of their electron configuration and bonding properties. Some nonmetals have such a strong tendency to form covalent bonds that atoms of a single element bond with one another in their non-molecular state. In other words, when the element is not in a compound, atoms of that element bond to themselves. Two typical examples are Cl and Br, both of which form one covalent bond with their single unpaired electron. Diatomic atoms satisfy the octet rule by sharing their single electron with another atom of the same element, which means that these elements appear as molecules until they react with other substances. The diatomic molecules are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2.

Ionic bonding

The transfer of electrons from one atom to another is called ionic bonding. Atoms that lose or gain electrons are referred to as ions. The gain or loss of electrons will result in an ion having a positive or negative charge. Here is an example: Take an atom of sodium (Na) and an atom of chlorine (Cl). The sodium atom has a total of 11 electrons as well as one electron in its outer shell. The chlorine has 17 electrons as well as 7 electrons in its outer shell. From this, the atomic number, or number of protons, of sodium can be calculated as 11 because the number of protons equals the number of electrons in an atom. When sodium chloride (NaCl) is formed, one electron from sodium transfers to chlorine. Ions have charges. They are written with a plus (+) or minus (-) symbol. Ions in a compound are attracted to each other because they have opposite charges.

Covalent bonds

A covalent bond is formed when two electrons are shared by two atoms. For example, H2 is formed as:

$$\mathrm{H}\bullet + \bullet\mathrm{H} \longrightarrow \mathrm{H}{:}\mathrm{H}$$

Each electron in the shared pair is attracted to each of the two nuclei participating in the bond. It is this attraction that holds the hydrogen atoms together and is responsible for covalent bonding in this and other types of molecules. Only valence electrons are involved in the covalent bonding of atoms with many electrons, i.e., atoms other than hydrogen. The octet rule provides a means of accounting for the valence electrons in a molecule. It states that for atoms other than hydrogen, bonds tend to form until the individual atoms are surrounded by eight valence electrons.

Covalent bonding is characterized by the sharing of one or more pairs of electrons between two atoms or between an atom and another covalent bond. This produces an attraction to repulsion stability that holds these molecules together. Atoms have the tendency to share electrons with each other so that all outer electron shells are filled. The resultant bonds are always stronger than the intermolecular hydrogen bond and are similar in strength to ionic bonds. Covalent bonding occurs most frequently between atoms with similar electronegativities. Nonmetals are more likely to form covalent bonds than metals since it is more difficult for nonmetals to liberate an electron, electron sharing takes place when one species encountered another species with similar electronegativity. Covalent bonding of metals is important in both process chemistry and industrial catalysis.

There are two types of covalent bonds: polar and non-polar. Polar bonds in molecules depend on the difference in electronegativity between the individual atoms in the compound and the compounds degree of asymmetry resulting in the unequal sharing of the electrons. A completely polar bond is more correctly described as an ionic bond and occurs when the difference in the electronegativity between 2 atoms is so great that one atom takes the electron of the other. If the difference in the electronegativity between 2 atoms of a molecule is between 0.4 and 1.7, the bond is polar. Non-polar bonds in molecules occur when the electronegativity of the individual atoms is relatively equal, resulting in the equal sharing of electrons. Most non-polar molecules are hydrophobic at room temperature. The most common form is the hydrogen bond.

The bond order indicates how many pairs of electrons are shared between atoms forming a covalent bond:

- The single bond is the most common type of covalent bond. It shares one pair of electron between two atoms.
- All other bonds with more than one shared pair of electrons are considered multiple covalent bonds.
- The double bond exists when 2 pairs of electrons are shared. An example is ethylene in which the sharing is between the carbon atoms.
- The triple bond exists when three pairs of electrons are shared. An example is hydrogen cyanide in which the sharing is between Carbon and Nitrogen.
- Quadruple bonds exist although they are rare and are mainly found in transition metals.
- Quintuple bonds have been found to exist in some dichromium compounds.
- Sextuple bonds can exist in diatomic dimolybdenum and ditungsten.
- Other bonds that are rarer include molecules with three center bonds.

Valence bonds

The model for valence bonds has been supplemented in recent times with the molecular orbital model. This model illustrates the atoms being brought together and the formation of hybrid molecular orbitals by an interaction with atomic orbitals. The molecular orbitals are a cross between original atomic orbitals and usually will extend between two bonding atoms. Quantum mechanics, which is the modern theory of action that is applied on very small scales, can be used to calculate the electronic structure, bond angles, bond distances and energy levels of simple molecules with great accuracy. Bond distances can be determined with as much accuracy as measurement provides. Energy calculations for small molecules are accurate enough to be used in determining thermodynamic heats of formation.

Molecular geometry

Bond angle, bond length, and conjugation

Bond angle is the angle in degrees between two bonds of the same atom. Imagine axes running from the nuclei of bonded atoms to the central atom's nucleus. The angle between two such axes is the bond angle. It is a fundamental variable in molecular geometry and is influenced by specific atomic properties such as electronegativity and the number of electrons. Bond length is another fundamental variable; it is the distance between the nuclei of two bonded atoms. It increases with the size of the bonded atoms, and is predictable based on periodic trends. Bond length between atoms of the same element decreases as the number of electrons involved in the bond increases; a double bond is shorter than a single bond, and a triple bond is shorter than a double bond. Conjugation is the alternation of single and double bonds within a compound; the varying bond lengths influence the overall geometry.

Significance to chemistry

The consistent ways in which atoms combine to form compounds enable chemists to predict the shapes, or geometries, of the molecules of any given compound. Knowing the shapes of molecules can in turn allow researchers to better understand the nature of matter such as crystalline structures, the nature of chemical reactions and how to facilitate or impede them to reach a desired outcome, such as in developing new drugs or petroleum refinement technology. On a molecular level, the shape of a molecule can increase or decrease its reactivity because the electrons available for reactions are less accessible to other particles in more complexly shaped molecules. If part of a molecule blocks its reactive site(s), a reaction is less likely to occur, and sometimes a molecule of a specific shape requires another specific shape to react with it.

VSEPR theory

The Valence Shell Electron Pair Repulsion theory is based on the fact that particles of similar charges repel one another. Electrons surrounding an atom will therefore arrange themselves to be as far apart from one another as possible. Using these laws, VSEPR theory predicts how electron pairs will arrange themselves around a given nucleus, making important predictions about the reactivity and geometry of molecules and atoms. The Lewis structure for barium, :Ba: , shows the electron pairs at their maximum possible separation of 180°. Boron, containing three valence pairs, forms bonds at 120° angles because its electron pairs must be evenly spaced on a 360° plane. All pairs are not equal; bonded pairs of electrons exert less repulsion than nonbonded pairs. VSEPR theory thus predicts an angular shape for H20 molecules: the nonbonded electrons in O press the H atoms closer together than the 180° one might expect.

Intermolecular forces

Intermolecular forces, also called van der Waals forces for their 19th century Dutch discoverer, are the electrostatic (or ionic) attractions between molecules. Because molecules in gases move so fast and are so far apart, the intermolecular attraction in gases is not an important factor in reactions. These forces are strongest in liquids and solids where they may determine the alignment of the molecules. Dipole interactions occur when the negatively and positively charged poles of molecules are attracted to one another; if the attractions are strong enough, the molecules form loose bonds, with the resulting structure shaped by the way the poles attract one another. Tightly bound molecules form solids, while loosely bound molecules form liquids. Nonpolar intermolecular attractions may occur when the electrons within a molecule are momentarily attracted to the positively charged nucleus of another molecule.

Coulomb force

There are two types of electric charge, positive and negative. When a negative charge has been rubbed off a cat's fur onto one's hands, the cat is left positively charged and the hand negatively charged. At the basis of this is the idea that action at a distance is caused by charges which are the sources of forces. With the electric "Coulomb force" the magnitude is in proportion to the product of the charges and inversely proportional to the square of the distance between them. This vector can be written either:

$F x = q 1 q 2 R x / 4 \prod \varepsilon \boxed{?} r^{3}$,
or
$F = q 1 q 2 / 4 \boxed{?} \prod \varepsilon r^{2}$.

In the equation, q1 and q2 are the charge magnitudes. R is a vector pointing from one to the other and r acts as the distance in between them.

London forces and hydrogen bonding

In nonpolar molecules, electrons are normally distributed evenly throughout their respective electron clouds, or energy levels. But those molecules can become temporarily polarized, having regions in which electrical charges are stronger or weaker than nearby regions. In this polarized state, the molecules are attracted to one another through weak forces called London forces.

The hydrogen bond is the attractive force between an H atom in one molecule and any electronegative atom in another nearby molecule. This bond is very prevalent in nature and is weaker than covalent or ionic bonds, but stronger than other intermolecular forces. Water molecules are highly polar and the H atoms of water molecules are strongly attracted to the O atoms of other water molecules. Similar H bonds are operable in NH3 and HF. These bonds are represented in structural formulas by a series of four dots between H and an electronegative atom of a neighboring molecule.

IUPAC nomenclature and common names

The International Union of Pure and Applied Chemistry developed and recommends a set of conventions and standards for naming chemical compounds. Collectively, these conventions are known as "nomenclature." The aim of using standardized nomenclature is to eliminate ambiguity or confusion over the composition of a given substance under discussion. In the ideal application of the conventions, the correct chemical formula of a substance can be derived from its name and its structure can often be described by proper application of the standards. Nomenclature uses prefixes, suffixes, and infixes to indicate the formula for and, if necessary, the placement of constituent parts within compounds. Many compounds were named before the standardization of nomenclature guidelines and so do not conform to them. Names such as formaldehyde, propane, and methane, for example, do not identify the numbers of C atoms present as IUPAC rules dictate. These are referred to as common names.

<u>Inorganic binary compounds</u>

The standard steps for naming inorganic binary compounds are:

- When a metal is present, name the metal first.
- Name the root of the non-metal, and add the suffix "-ide."
- Examples: sodium chloride, aluminum bromide, potassium sulfide, sodium pentoxide.
- When a metal is not present, the first element to be named is determined by starting at the left of the periodic table; the left-most element is named first. A Greek prefix must indicate the number of the prefixed atom present in the molecule. When only one atom of the first element is present, no prefix is necessary.
- Examples: carbon monoxide, dinitrogen tetrachloride, triphosphorous pentafluoride. Hydrogen dioxide.
- When properly followed, conventions of nomenclature describe the quantities and nature of chemicals in the compounds named.

<u>Acids</u>

Because they form an important class of chemicals, acids have their own slightly different system of nomenclature. Hydrogen-based acids have the prefix "hydro," followed by the root of the halogen, and ending with the suffix "-ic." Thus, HF, "Hydrogen fluoride," is called in solution "Hydrofluoric acid." Other examples include: HCl – hydrochloric acid, HBr – hydrobromic acid, HI – hydroiodic acid. Another class of acids is very abundant in nature: oxygen based acids—oxoacids, or carboxylic acids—contain hydrogen, oxygen, and another element. Many of them have common names, but their IUPAC names are derived by naming the oxoanion, using the suffix "-ic," and adding acid. Thus NO3-, a nitrate anion, combines in solution with hydrogen and is called nitric acid—HNO3. "Sulfate" becomes sulfuric acid; "chlorate" becomes "chloric acid," and so on.

Organic compounds

While the naming of inorganic compounds depends on the type of bonds between molecules, nomenclature for organic compounds is most directly driven by the functional groups present within it. Organic nomenclature describes the bonding between atoms as rings and/or chains, and identifies the type and location of functional groups within the compound. The names of organic compounds will generally have three parts: a root indicating the largest chain or ring structure of the carbon atoms present, a suffix designating the molecule's functional group(s), and names of other groups or elements present. Other rules for organic nomenclature vary depending on the type of hydrocarbon chain forming the parent, or root, and the types of bonds present within that chain. Cyclic hydrocarbons, those formed in rings, also have slight variations in nomenclature conventions applied to them.

In naming organic compounds, the longest carbon chain or ring, also called the parent hydrocarbon, forms the root of the compound name. These roots correspond to the number of C atoms present in the longest chain and are classified into groups such as alkanes, alkenes, and alkynes. Alkenes contain only saturated hydrocarbons in which the C atoms are bound with single bonds. Alkenes are unsaturated chains containing double bonds, and alkynes are unsaturated chains in which the C atoms are joined in triple bonds. Roots may belong to any of these classes as indicated by their suffix. The following are some common parent (or root) names and their corresponding number of C atoms:

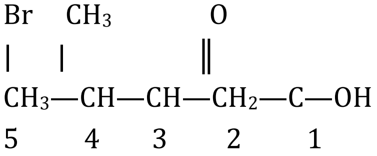
Root name	# of C atoms	Root name	# of C atoms
Meth	1	Undec	11
Eth	2	Dodec	12
Prop	3	Tridec	13
But	4	Tetradec	14
Pent	5	Icos	15
Hex	6	Henicos	16
Hept	7	Docos	17
Oct	8	Triacont	18
Non	9	Tetracont	19
Dec	10	Pentacont	20

According to IUPAC nomenclature guidelines, suffixes describe the family group or functional group/groups present in the named compound. Each functional group is associated with a corresponding suffix which is to be attached to the root name of the parent hydrocarbon. Some common functional groups and their suffixes are as follows:

Group	Suffix	Group	Suffix
Alkane	-ane	Ketone	-one
Alkene	-ene	Carboxylic acid	-oic acid
Alkyne	-yne	Alcohol	-ol
Aldehyde	-al	Ether	-ether
Sulfide	-sulfide	Amine	-amine
Ester	-oate	Amide	-amide

A compound including a propanoate, therefore, is formed from an alkane, "prop," parent hydrocarbon (CH3CH2CH3) with an ester functional group attached. There are actually two suffixes in propanoate: -ane indicates an alkane (a saturated non-cyclic hydrocarbon) and –oate indicates an ester.

Names of organic compounds, which can be very complex molecules, use numbers to indicate the location of groups or substituents attached to the carbon chain. To determine these numbers, each C atom or CH molecule in the ring or chain is assigned a number designating its proximity to the highest priority functional group. The example below shows an organic compound with a carboxyl group attached at the far right. The C atoms are numbered starting with 1 at the carboxyl C atom, and moving to the left. Thus the C atom on the far left is the 5th C atom in the parent chain. The IUPAC name indicates that the Br is attached to the 4th C atom and the methyl group ($CH3$) is attached to the 3rd.



$$\begin{array}{ccccccccccc} Br & & CH_3 & & & & & & O & & \\ | & & | & & & & & & \| & & \\ CH_3 & — & CH & — & CH & — & CH_2 & — & C & — & OH \\ 5 & & 4 & & 3 & & 2 & & 1 & & \end{array}$$

4-bromo-3-methylpentanoic acid

Alkanes

Alkanes present unique challenges for nomenclature because of the many isomers (structural variations) possible. The standard IUPAC conventions for organic compounds apply: each name consists of a prefix, a parent (or root), and a suffix. But the names of alkanes further specify the length of the C parent and the position of any alkyl groups (an alkane which has lost an H atom, becoming $CH3$). The prefix indicates the type and location of alkyl groups, but a new suffix, --yl, indicates the presence of an alkyl. The simplest alkane, for instance, is methane ($CH4$); when an H is lost, it becomes methyl ($CH3$). Alkyl groups are generally represented in structural or condensed structural formulas as R—.

Alkenes

The IUPAC rules for naming alkenes are very similar to those for alkanes, but because the double bond present in the chain is the primary identifying characteristic of an alkene, the location of the bond must be specified. The longest C chain containing the double bond is the parent chain and a Greek prefix identifies the number of C atoms. An alkene with eight carbon atoms is therefore an octene (a chain of eight C atoms with no double bonds is an octane). Counting C atoms from the end of the chain nearer the double bond, the number of the first C in the double bond precedes the parent name, as in 3-octene. Using the same numbering method, the positions of alkyls or other groups in the chain are also identified.

Alkynes and aromatic compounds

The distinguishing characteristic of alkynes is the triple bond; therefore, the nomenclature conventions must identify the presence of the triple bond with the suffix –yne and its location with a prefixed number derived from the same numbering method as with alkenes. 3-octyne therefore indicates a hydrocarbon chain containing a CĐC triple bond which begins at the 3rd C atom from the end. Prefixed numbers also indicate the bond location of alkyl or other substituents groups.

Aromatic compounds are those in which a hydrocarbon ring, called a benzene ring ($C6H6$), is the parent. They are named by adding substituent groups as prefixes, for example: a simple compound in which bromium is attached to a benzene is $BrC6H6$, or bromobenzene; an ethyl group ($CH2CH3$) attached to benzene is called ethylbenzene; $NO2$ attached to a benzene is nitrobenzene, and so on.

Esters and ethers

Ethers are commonly named alkyl ethers because of their chemical structure (C—O—C). When other groups or alkyls are present, they are named in alphabetical order followed by "ether," as in butyl methyl ether. Formal IUPAC guidelines, however, dictate that ethers be called alkoxyalkanes in which the smaller alkyl and the O atom form an alkoxy group. The ether, or alkoxy group, is considered a substituent on the longest parent chain or ring. For example, a six carbon alkane is a hexane; if an alkoxy group is attached to the hexane at the fourth carbon atom, the compound is 4-methoxyhexane. For esters, the alkyl group bonded to the O is named first, followed by the parent root with the –ate suffix indicating an ether, as in methyl benzoate ($OCOCH3$).

Aldehydes and ketones

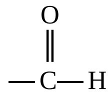
Familiar aldehydes and ketones have common names such as formaldehyde and acetone, but the IUPAC conventions apply to more complex compounds. Aldehydes are named by identifying the longest C chain containing the C=O bond. Each substituent group is then named, preceded by the number of the C atom to which it is attached; the count of C atoms begins at the C=O bond. The suffix –al identifies the compound as an aldehyde, as in 3-hydroxy-2-methylbutanal. The C=O bond is understood to be on the terminal C atom of the chain. Ketone nomenclature differs in that the C=O bond is not the terminal atom, so it must be numbered in a prefix to the name of the compound. Count the C atoms beginning at the end nearest the double bond. A suffix –one indicates a ketone, as in 4-methyl-2-pentanone.

Pauling scale

Acclaimed chemist Linus Pauling developed the Pauling scale in 1932. On the scale, an electronegative value of 3.98 is given to the most common electronegative chemical element, which is fluorine. Texts may often use the value 4.0. The least electronegative element, which is francium, has a value of 0.7. The rest of the elements have values that are in between these. Hydrogen is arbitrarily given a value of 2.1 or 2.2 on the Pauling scale. Bonds between atoms with a large electronegative difference, those more than or equal to 1.7, are normally considered to be ionic. Those with values between 1.7 and 0.4 are normally thought of as polar covalent. Those with values below 0.4 are non-polar covalent bonds. An electronegativity difference of 0 indicates a covalent bond that is entirely non-polar.

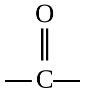
Organic chemistry and biochemistry

Organic chemistry is the study of organic chemicals, all of which contain carbon. These compounds are usually associated with plants, animals, and other life forms; while inorganic compounds are associated with minerals and do not usually contain carbon. Biochemistry seeks to understand the mostly organic compounds involved in life processes and how they interact and change to support life functions. 99% of the compounds necessary for life are organic, but some are mineral atoms, such as phosphorus, potassium, iron, chlorine, magnesium, and calcium. Biochemistry, therefore, encapsulates organic chemistry, but an organic chemist need not know a great deal about biochemistry. Many millions of compounds are studied and described by these two sub fields of chemistry.



$$-\overset{\overset{\displaystyle O}{\|}}{C}-H$$

Aldehyde



$$-\overset{\overset{\displaystyle O}{\|}}{C}-$$

Ketone

Carbohydrates

Carbohydrates are important organic molecules containing carbon, hydrogen, and oxygen; they are the primary source of energy for most living organisms. Carbohydrate digestion is the result of several fundamental metabolic pathways and each process involves a great many subsidiary reactions. The formation of glucose from carbohydrates is one important pathway. Glucose ($C_6H_{12}O_6$) is one of several monosaccharides produced from carbohydrates by chemical reactions. Glucose molecules then enter another metabolic pathway called glycolysis in which they are broken down into pyruvate ($C_3H_4O_3$), which in turn begins another series of cascading reactions which lead to usable energy. The citric acid cycle is another important metabolic pathway in the digestion of carbohydrates.

Lipids

Lipids are bimolecular that are insoluble in water but soluble in other nonpolar solvents. They are responsible for the waxy or greasy nature of substances found in plants and animals. Saponification is a process used to further distinguish classes of lipids; it is the process by which esters are hydrolyzed in base solution to yield glycerol and the salts of fatty acids. If lipids do not contain esters, they are nonsaponifiable and are classified accordingly. Hydrolysis is another important reaction for lipids. In the presence of water and an acid catalyst, lipids break down into glycerol and fatty acids. Hydrogenation is the process of converting C=H double bonds to C—H single bonds resulting in increased saturation of the hydrocarbon and a correspondingly higher melting point.

Enzymes

Some of the most important functions associated with enzymes regulate their activity. Enzymes can exist in an inactive or preactive state called zymogens or proenzymes. They are stored until needed by a cell for a specific reaction and are then released and activated to become active enzymes. They are activated by breaking one or more peptide bonds on the enzyme through hydrolysis reactions. Enzyme activity is also regulated by substances called modulators which bond to the enzyme and change its three dimensional orientation, rendering it temporarily inactive. Modulators can be either activators, which increase the enzyme's activity, or inhibitors, which decrease it. Some substances act as uncontrollable, irreversible inhibitors and are thus highly toxic. Cyanide and other heavy metals are deadly because of their inhibitory functions. Enzyme inhibition is a process in which products at the end of reactions inhibit earlier reaction steps, effectively regulating the concentration of target compounds.

Enzymes are catalysts in metabolic pathways and play roles in very large numbers of reactions; all of them are globular proteins and catalyze reactions involving other proteins, making them of fundamental importance in complex bioprocesses. Most enzymes are highly selective, catalyzing only specific reactions when the reactants come into contact with the reactive site of the much larger enzyme molecule. For instance, the enzyme urease is a catalyst in only one reaction: the hydrolysis of urea. By contrast, other catalysts, such as strong acids or bases, catalyze they hydrolysis of any amide. Enzymes are complex proteins that are highly catalytic, increasing the rate of some reactions by as much as 10^{20} times. Enzymes are named systematically with a set of conventions much like IUPAC nomenclature. These conventions specify the reactant and functional group acted upon and the type of reaction catalyzed. A suffix –ase indicates the named compound is an enzyme.

Nucleic acids

Nucleic acids are polymers composed of many smaller nucleotides which can be further broken down to phosphoric acid, organic bases, amines, and deoxyribose or ribose. The nucleotides link to form the polymers in an alternating sequence of phosphate units and sugar units. A nucleic acid containing deoxyribose as its sugar is deoxyribonucleic acid, or DNA. If ribose is the sugar, the compound is ribonucleic acid, or RNA. Nucleic acids are important substances in biochemistry; DNA is the fundamental constituent of genes carrying hereditary information and occurs in a complex molecular arrangement known as the double helix. RNA is a compound essential to the proper formation and utilization of proteins within cells and is found in several varieties including the most common tRNA (transfer RNA) and mRNA (messenger RNA).

Codons

Codons are long molecules formed of nucleotides and found in mRNA. Each codon is specific to the protein in whose synthesis it plays a lead role; each one has unique properties such as molecular weight. Codons are long because they contain three nucleotides for each amino acid required by their target protein. They are attached to ribosomes, which can move through several codons of the mRNA during protein synthesis. The ribosome chemically activates the appropriate codon to form amino acids in proper sequence for the necessary protein. It is a codon which eventually signals the completion of the protein and the separation of the peptide chain from the ribosome.

Carbon

Organic chemistry is to a large degree dependent on the properties of carbon. All organic compounds contain C (although some inorganic compounds also contain C); this is the basis of all life on Earth. The primary property of C responsible for its importance is its ability to form very stable covalent bonds with other C atoms or with other elements. It has four valence electrons, making it an important partner with other atoms according to the octet rule. Another important characteristic of C is its ability to form many different isomers in compounds with other elements, particularly H and the functional groups. Finally, C is important because of its tendency to hybridize, or mix s and p orbitals to form a sp3 hybrid orbital. Carbon's electron configuration by itself does not adequately explain its tetravalent bonding ability; hybridization is necessary to explain it.

<u>Primary, secondary, and tertiary carbon</u>

Primary, secondary, and tertiary carbon atoms are those in a structure or molecule bonded to only one, two, or three other C atoms, respectively. These terms are useful for the specification of C atoms in organic compounds. Because C is involved in so many complex molecules, it is important to be able to distinguish one C atom from another; these designations are one way of doing so. Knowing which type of C atom is under consideration can give clues about its position in a chain or ring. A primary C, for instance, may be the end of a hydrocarbon chain and may have other atoms or groups of atoms attached; a primary carbon may also be part of a double or triple bond (i.e., an alkene or alkyne).

<u>Carbocation, carbanion, carbene, and carbenium ion</u>

Because of Carbon's foundational importance to nomenclature and organic chemistry in general, names were devised to specify slight variations in the structure and charge of C atoms when bonded under circumstances which give them unique characteristics. A carbocation is a positively charged C ion; they occur under various circumstances and may be part of much larger molecules. A carbanion is a negatively charged C ion which features an unshared electron pair on a carbon atom, such as

H_3C-. Carbene is a general name for any molecule H_2C:, or molecules formed from substitution reactions within that configuration. The molecule is electrically neutral and contains a nonbonding electron pair on the C atom. Carbenium ions are carbocations having at least one structure attached which contains a C atom with a vacant p orbital and hence a slight positive charge.

Hydrocarbons

Hydrocarbons are molecules containing only C and H and form the basis of organic chemistry. They bond together with strong covalent bonds in chains or rings to form the backbone of organic molecules which may have any number of a large variety of functional groups attached. Hydrocarbons are classified into two large groups based on the bonds between their C atoms; if only C—C single bonds are present, the hydrocarbon is saturated because other electrons will then form bonds with surrounding H atoms. If the C atoms are bound with multiple bonds (C=C or CĐC), fewer electrons are available to form bonds with H atoms and the hydrocarbon is unsaturated. Hydrocarbons are very stable molecules because of the tetravalency of C; it has four valence electrons making its octet requirement easy to satisfy, which results in very strong bonds.

Organic compounds

There are approximately 9 million known organic compounds, each with its own chemical and physical properties. Listing them all and attempting to classify them by their properties would thus be a meaninglessly complex endeavor. But organic compounds are known to be composed of a hydrocarbon backbone and one or more functional groups attached to the hydrocarbon framework. The functional groups are the basis for organic nomenclature and have characteristic traits and predictable reaction behavior which strongly influence the molecule to which they are attached. Separating the backbone from the functional groups provides a more manageable way to classify large molecules in large numbers because the molecule behaves and reacts according to its functional group since they are the sites of reactions in organic molecules.

Classes

Heterocyclic compounds are organic compounds in which one or more atoms other than C have combined with C to form a ring structure. Parent hydrocarbons may form rings and are then named by adding “cyclo” as a prefix as in cycloalkane and cycloalkene. These are usually the result of substitution reactions; for example, a N atom takes the place of a C atom in a cycloalkane. Aromatic compounds are those whose parent is a benzene ring—a cyclic C compound containing 6 C atoms. Aliphatic compounds are those whose parent hydrocarbons take the form of a chain. Biomolecules are large, complex organic molecules that are essential to biological processes; examples are lipids, carbohydrates, enzymes, proteins, etc.

Physical properties

Organic compounds are usually covalently bonded (though specific molecules may be polarized) and, therefore, are less soluble in water than inorganic compounds whose ionic bonds are easily separated by polarized water molecules. The intramolecular forces in organic molecules are relatively weak compared to the same forces in inorganic compounds; this is also due to the prevalence of covalent bonds. Organic compounds' melting and boiling points are below 300° C, while inorganic melting and boiling points can be much higher. They react more readily with oxygen in combustion (though activation energy is required) and are poor conductors of electrical current. They occur in nature more frequently in gaseous and liquid states, or as solids with low melting points. Inorganic compounds, by contrast, more often appear as solids with high melting points.

Energetics

Joule

The joule, expressed as "J," is the SI unit of energy or work. It was named after physicist James Joule. It is a derived unit that reflects the work done or energy required to exert a force of one newton for a distance of one meter. The same quantity could also be referred to as a newton meter with a symbol N-m. The newton meter is normally utilized to measure torque instead of energy. One joule is the absolute minimum amount of energy that is required on the Earth's surface to lift a one-kilogram object 10cm. It is the equivalent of about 2.390x10-4 calorie or 9.48x10-4 BTU.

Heat and temperature

Temperature is the measure of the average energy of molecular motion within a substance. Heat is the total energy of molecular motion. Heat energy is dependent upon the speed of particles, the size or mass of the particles and what kind of particles are in an object. Temperature is not dependent upon types of objects or their size. Heat increases or decreases the temperature. If heat is added the temperature is higher. If heat is removed, then the temperature drops. Those higher temperatures signify molecular movement as well as vibration and rotation with more energy. Heat is energy while temperature is not energy but is instead a measure of it.

Thermocouple

A *thermocouple* is a type of temperature sensor that may also be used to convert thermal energy into electrical energy. This device can measure wide ranges of temperature, but is prone to inaccuracy. The thermocouple works on the principle that when any conductor (such as a metal) is subject to a thermal gradient, it will generate a small amount of voltage. This is called the Peltier-Seebeck effect, and is a direct conversion of heat differentials to electric voltage, or the reverse. In order for a thermocouple to produce an output voltage, there must be a temperature difference between the junction of two metal wires.

Temperature

Temperature is a thermodynamic property of a substance. The temperature of a substance depends upon its energy content. The two most commonly employed temperature scales are Fahrenheit and Celsius. In the Fahrenheit scale, water freezes at 32°; in the Celsius scale, it freezes at 0°. There are also two absolute scales of temperature, which are used to define the temperature independent of the properties of the substance. The absolute scale in the U.S. system is the Rankine scale, and the absolute scale in the SI system is the Kelvin scale.

Specific heat of metal

Measurement of the specific heat of metals such as copper can be done by placing the metal into a calorimeter that contains water at a different temperature than the metal. If the copper is at first a higher temperature than the water, the heat will transfer from the copper to the water. While the metal is cooling, the water will warm. As both reach similar temperatures no more heat transfers. The calorimeter minimizes heat loss to surroundings. Heat that is gained by the water will be equal to the heat lost by the copper. The mass and initial temperature of both the water in the calorimeter and the copper must be measured first. The copper is put into boiling water long enough to reach the temperature of the water and then is put quickly into the calorimeter. Once water and copper reach equal temperatures, the final temperature reading is taken.

Absolute zero

Absolute zero is the lowest possible temperature that is obtained when a system is at its minimum possible energy. The Kelvin scale has its zero point at absolute zero which is -273.15 on the Celsius scale and -459.67 Fahrenheit. A true minimum has been confirmed by a number of experiments. Since temperature is molecular energy then it would follow that there must be a certain point at which no additional energy could be taken from a system. It is possible to approach absolute zero, yet the Third Law of Thermodynamics says that absolute zero cannot be reached in a system. Cosmic background radiation, which are relic photos left over from the very early and hot phase of the Big Bang, has a temperature of about 2.7 degrees above absolute zero.

Calorimeter

A calorimeter is an object that is used for measuring the heat of chemical reactions or physical changes. A calorimeter might be composed of just a thermometer attached to an insulated container. In order to find the enthalpy change per mole of a substance X in a reaction between the liquids X and Y, they are added to the calorimeter and the first and last temperatures are recorded. A value for the energy given off during a reaction, if that reaction was exothermic, can be found by multiplying the temperature change by the mass and specific heat capabilities of the liquid. The enthalpy change of reaction can be found by dividing the energy change by the number of moles X present.

Mercury-in-glass thermometer

The mercury-in-glass thermometer uses mercury in a glass tube. It contains marks on the tube that are calibrated and allow the temperature to be seen by the length of the mercury inside the tube. That length will vary according to temperature. A bulb is usually at the base of the thermometer in order to increase the sensitivity of the instrument. The bulb contains most of the mercury. The expansion and contraction of the volume of mercury is amplified in the bore of the tube, which is much more narrow than the bulb. The space that is above the mercury may be a vacuum or it might be filled with nitrogen. Mercury freezes at -38.83 degrees C or 37.89 degrees F. Therefore, it can only be used at temperatures higher than 37.89 degrees F. Since mercury does not expand when frozen, like water, it will not break the glass tube.

Heat conduction

Conduction is one method of heat transfer and happens when there is molecular agitation without motion of the material. For instance, if one end of a metal rod has a higher temperature, then energy will transfer down that rod toward the colder portion because higher speed particles will connect with slower ones that will result in a net transfer of energy to the slower ones. In the heat transfer between two plane surfaces, such as in heat loss through the wall of a home, the conduction heat transfer rate variables are: Q = heat transferred in time = t; k = thermal conductivity of the barrier; A = area; T = temperature; d = barrier thickness.

The rate of the heat transfer is dependent upon temperature gradient and thermal conductivity of a specific material.

Heat convection

Convection is another heat transfer method. The heat is carried by the mass motion of a fluid such as water when the heated fluid is moved away from the heat source and carries energy along with it. Convection occurs above a hot surface because hot air expands, becomes less dense and also rises. Hot water is less dense than cold water and rises. This causes convection currents transporting energy. Convection can also cause circulating liquid as in a heated pot of water over a fire. Heated water will expand and become more buoyant. The more dense, cooler water near the surface descends and circulation patterns may form. Convection is believed to be a major factor in the transportation of energy from the center of the Sun to its surface.

Misunderstandings about heat energy

Students may believe that heat and temperature are the same thing. They may think that heat and cold are similar quantities but opposite. They may believe ice cannot be cooled below freezing. Their beliefs may also include the thought that the temperature of water cannot be higher than its boiling point. They may think that the temperature always rises when a mixture is heated. Or they could think that heat rises and energy may only go up. They may mistakenly think that bubbles in boiling water are air or that you can tell the temperature of an object by the way it feels. Students may also hold views that while matter contains particles it does not consist exclusively of particles. They may think particles are all the same size and about the size of cells or perhaps a bit smaller.

Gibbs' Phase Rule

The phase rule, postulated by J.W. Gibbs in the 1870s, is a description of the possible number of degrees of freedom in a closed system that is at equilibrium.

The rule is: $F = C - P + 2$

Description of Variables:

- F = The degrees of freedom. This is the number of intensive variables and those that are independent of the quantity of material present that need values specified to fully determine the state of the system. Variables might include pressure, concentration or temperature.
- P = Phase. The phase is a component of the system that cannot be mixed or blended with other parts such as a gas, liquid or solid. A phase may contain several chemical constituents. Those may or may not be shared with other phases.
- C = Chemical Constituents. These are the distinct compounds, or elements, that are part of the equation of the system. If some of the constituents stay in equilibrium with each other, regardless of the state of the system, then they should be considered a single constituent.

Graham's Laws of Diffusion

Some physical properties of gases are dependent upon their identity as a gas. Thomas Graham, in 1829, used an apparatus to study the diffusion of gases, or the rate at which two gases will mix. Using a similar apparatus can exemplify this diffusion. By taking an apparatus consisting of a glass tube that is sealed at one end with plaster and has holes large enough to allow gas to enter or leave the tube and filling it with Hydrogen gas, the water level rises slowly because the hydrogen gas molecules escape through the holes in the plaster faster than the molecules in the air can enter the tube. Graham was able to obtain information on the rate at which gases mix by studying the rate at which water levels rose. Graham's law states that the rates at which gases diffuse is inversely proportional to the square root of their densities.

Graham's Law of Effusion

Thomas Graham studied the rate of gas effusion, which is the rate at which gas will escape through a pinhole into a vacuum. The rate of gas effusion is inversely proportional to the square root of either the gas density or the molecular weight of the gas. To show this: A filter flask is evacuated with a vacuum pump. A syringe is filled with 25 mL of gas. The time required for the gas to escape through the syringe needle and into the evacuated filter flask is measured with a stop watch. The results can show that the time required for the 25-mL samples of different gases to escape is proportional to the square root of the molecular weight. Graham's observations about the rate at which gases mix, or diffuse, or effuse suggest that particles of gas which are relatively light such as H2 or He move faster than heavier gas particles such as CO2 or SO2.

Avogadro's law of gases

Avogadro's law relates the volume and number of molecules of a gas. It holds that equal volumes of different gases contain the same number of molecules. The volume of any given gas is proportional to the number of molecules present. This is a minor aspect of the law:

$$\frac{V}{n} = a$$

where V is the volume of the gas; n is the number of moles in the gas; and a is a constant. Most important is that the ideal gas constant has the same value for all gases. This means:

$$\frac{p_1 \cdot V_1}{T_1 \cdot n_1} = \frac{p_2 \cdot V_2}{T_2 \cdot n_2} = const$$

has the same value for all gases regardless of the gas molecule size.

Ideal gas

A gas can be considered ideal if its pressure is very low and its temperature is higher than its critical temperature. This allows the molecular radius to be considered insignificant compared to the distance between molecules. It can also be assumed that the molecules do not come into contact with each other. Typically, highly superheated vapors are considered ideal gases. Avogadro's law states that equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules. This is represented for n moles of any gas as the equation of state for ideal gases, or the ideal gas law:

$$PV = nRT$$

Where P = pressure, V = volume, n = number of moles, R = universal gas constant, T = temperature.

Boyle's law of gases

Boyle's law relates the pressure and volume of a gas. It states that the volume of a gas and its pressure are inversely related:

$$PV = k$$

where V is volume; P is pressure; and k is constant. Given a fixed amount of gas at a constant temperature, the volume is inversely proportional to the gas pressure. The volume and pressure are usually related by a constant, k1, called the proportionality constant. This says that the product of the pressure and volume of a fixed amount of gas at constant temperature is constant.

Charles' law of gases

Charles' law relates the temperature and volume of a gas. It states that the volume of a gas and its temperature are directly proportional:

$$\frac{V}{T} = k$$

Where V = volume; T = temperature; and k is a constant. Given a fixed amount of gas at a constant pressure, the volume is directly proportional to the absolute temperature of the gas. The volume and temperature are often related by a constant, k2, called the proportionality constant:

Pressure-volume relationship with gas

Gas molecules that are inside a fixed volume, such as a balloon, are constantly moving around freely. During this molecular motion, they often collide with each other and with the surface of any enclosure that might be there. The force of impact of a single collision is too minuscule to be sensed. But the large number of impacts of gas molecules will exert a sizable force onto the enclosure's surface. The larger the number of collisions per area of enclosure, the larger the pressure. The SI unit of pressure is Pa, or Pascal. In meteorology the use of millibars (mb) is accepted where 100 kPa = 1000 millibars. The direction of this gas pressure force is always perpendicular to the surface of the enclosure at all points.

Hydrostatic pressure

The principle that at a given level the pressure is equivalent to the weight of the overlying column is true for gases and liquids. The pressure generated by an overlying fluid column is called hydrostatic pressure. The air column's upper boundary that causes a rise in atmospheric pressure is the vacuum of space. Being somewhat light, the mass of an air column with a 1 cm2 cross section is almost exactly 1 kg. If a heavier liquid substance is used in balancing this air column, a relatively small length would be required. Additionally, the density of liquids does not change with height as most liquids are incompressible, so an equivalent liquid column has a well-defined upper boundary below a vacuum.

Compressibility of gas

When an object moves through a gas, the compressibility of the gas becomes an important factor. Gas molecules move around an object as it passes through. If an object passes by at a low speed, usually below 200 mph, the density of the fluid remains constant. For higher speeds, some of the energy of the object will go into compression of the fluid, moving molecules closer together and causing a change in the gas density. This alters the amount of force that results on the object. The effect becomes of more importance as speed is increased. Near and beyond the speed of sound, 700 mph, shock waves are produced. They affect both the lift and drag of objects.

Dalton's Law of Partial Pressure

Dalton's Law of Partial Pressure states that the total pressure that is exerted by the mixture of different gases in an gas volume is equal to the sum of the partial pressures of each individual component in the overall gas mixture. The law also states that each individual component in any mixture of gas exerts its own individual partial pressure, in the exact ratio as its mole fraction -- or volume-based concentration -- in the particular mixture. The relationship gives a way to determine a volume-based concentration of any individual gas components in a mixture of several other gases that are in air.

Enthalpy of phase changes

Enthalpy (symbol H and also known as heat content) is the sum of the internal energy of matter and the product of its volume multiplied by the pressure. The enthalpy of phase changes include vaporization, fusion and sublimation. The difference in enthalpy per mole of molecules between a gaseous and liquid state of a substance is known as the enthalpy of vaporization, or Hvap. The difference in enthalpy per mole of molecules between the liquid and solid states of a substance is known as the enthalpy of fusion, or Hfus. The difference in enthalpy per mole of molecules between the gaseous and solid states of a substance is called the enthalpy of sublimation, or Hsub.

Spontaneous processes

A spontaneous process is a chemical reaction in which a system releases free energy, usually heat, and goes to a lower and more thermodynamically stable state of energy. For reactions at constant temperatures and pressures where:

$$\Delta G = \Delta H - T\Delta S$$

- A negative ΔG depends on the sign changes in enthalpy (ΔH), entropy (ΔS) and the magnitude of the Kelvin absolute temperature.
- Changes of ΔG will not experience a direct change by temperature because it is never less than zero.
- When ΔS is positive and ΔH is negative, a spontaneous process occurs.
- When ΔS is positive and ΔH is positive, a process is spontaneous at high temperatures, where exothermicity has a small part in the balance.
- When ΔS is negative and ΔH is negative, a process is spontaneous at low temperatures, where exothermicity is a major factor.
- When ΔS is negative and ΔH is positive, a process is not spontaneous at any temperature, but there is a reverse spontaneous process.

Entropy

Entropy is a thermodynamic measurement of the amount of energy in a physical system that cannot be used to do work in a closed system. It is a key physical variable used in the description of a thermodynamic system. The SI unit of entropy is joules per kelvin; this is the same unit used for heat capacity. Entropy depends upon the current state of a system rather than its past history. In a process where the system gives up an energy ΔE, and its entropy falls by ΔS, a quantity at least TR ΔS of that energy must be given up to the system's surroundings as unusable heat (TR is the temperature of the system's external surroundings, which need not be the same as the system's current temperature T); Otherwise the process will not go forward.

Equipartition theorem of energy

The theorem of equipartition of energy states that molecules in thermal equilibrium have the same average energy associated with each independent degree of freedom of their motion. In thermodynamics the equipartition theorem says that the mean internal energy that is associated with each degree of freedom of a monatomic ideal gas is the same. For a molecule of gas, each component of velocity will have an associated kinetic energy.

Energy distribution function

The distribution function f(E) is the probability that a particular particle is in energy state E. The distribution function generalizes ideas of probability to where energy can be treated as a variable that is continuous. There are three distinct and different distribution functions: Maxwell-Boltzmann deals with particles that are identical but are also distinguishable. An example is the distribution of molecular speed. Bose-Einstein functions are directed at those particles that are indistinguishable with integer spin. Thermal radiation is an example. The Fermi-Dirac is concerned with identical and indistinguishable particles with a half-integer spin. An example would be electrons in a metal.

Maxwell-Boltzmann Distribution

The Maxwell-Boltzmann distribution is a classical distribution function for the distribution of a quantity of energy between identical particles that are distinguishable.
Certain classical statistical physicists surmise that:

- No restriction exists on the number of particles occupying a given state.
- The distribution of particles, at thermal equilibrium, among available energy states takes the most probable distribution that is consistent with the total available energy as well as the total amount of particles.

It is unlikely one particular particle will get an energy that is well above the average. Energies that are lower than average are more in favor because of more ways to get them. If a particle gets an energy of 10 times the average, then it will reduce the number of possibilities for the distribution of the remaining energy. It is thus unlikely because the probability of occupying a given state is proportional to the different ways it can be obtained.

Science of thermodynamics

Thermodynamics as a science investigates and analyzes thermal paths that lead from an energy source, such as a form of fuel, to useful work. The development of equipment that operates over thermal paths that are close to theoretical ones is sometimes viewed as an art. Nonetheless, the two are distinct and studying them separately help to give a better application of theoretical relationships to the problems involved in practical design. One reason is that theories can become obsolete more slowly than equipment. Also, a deep knowledge of thermodynamics is needed to relate to the equipment of today, yesterday and tomorrow. Thermodynamic study begins with the sources of energy and then leads to one of the many thermal paths that end up as the desired product or mechanical work. While a number of steps are required for conversion, each step is purposeful in liberation, storage, transportation, transformation, transfer or the utilization of energy.

Intensive and extensive properties

The thermodynamic state of a substance is determined by its properties. Intensive properties are independent of the quantity of material making up the system. Examples of intensive properties are temperature, pressure and stress. Extensive properties are dependent upon the quantity of material involved. Examples of extensive properties are volume, density, mass, charge, and strain.

Thermodynamic process

The thermodynamic process is defined as the series of continuous states that are followed by a working substance it liberates, transforms, transfers or receives energy from. A major problem in thermodynamics is classifying the different types of processes and picking the ones providing the best theoretical applications. Under constant conditions, three measurable properties exist. One is an isothermal process that occurs without a change in temperature. Another is an moist process which takes place at a constant pressure. There is also an isometric process which is a process of constant volume. An example of isothermal is when energy is added to ice and melting occurs. Isopiestic is represented by expansion against atmospheric pressure. Isometric processes are those that correspond to heating a material that is inside a rigid and nonexpanding container.

Independent thermodynamic properties and intensive state

The number of independent thermodynamic properties required to establish the intensive state of any system can be determined using the phase rule:

$$F = 2 - P + N$$

where P = number of phases, N = number of chemical species. For a pure homogeneous substance, P = 1, N = 1 therefore the number of independent thermodynamic properties, or degrees of freedom, F, that must be defined to fix its state is 2.

Phase rule variables are intensive properties and are independent of the extent of the system and the individual phases. As such, the information given by the phase rule is independent of both the size of the system and the amounts of the phases present.

First Law of Thermodynamics

The first law of thermodynamics in its simplest form holds that energy cannot be either created or destroyed. If energy enters a system it must leave the system or be stored in the system in some manner. Stated another way, the total quantity of energy is constant and when energy disappears in one form it simultaneously appears in another form. The work done in an adiabatic process depends only on the end conditions and not the nature of the process. The first law can be written as follows:

$$E = Q - W$$

Where Q is heat; E is energy; and W is work.

Second Law of Thermodynamics

The second law of thermodynamics restricts the direction in which heat and energy can flow. It has been stated in many ways, a very common and simple one is:

No process is possible which consists solely in the transfer of heat from one temperature level to a higher one. Stated another way, it is impossible to convert heat absorbed by a system completely into work done by the system. This means that the net entropy of a process will always increase and will proceed in the direction which causes the entropy of the system and surroundings to increase.

Third Law of Thermodynamics

The Third Law of Thermodynamics states that as a system approaches absolute zero of temperature all processes cease and the entropy of the system approaches a minimum value, or zero for the case of a perfect crystalline substance. The law states that the entropy of the system at absolute zero temperature is a constant that is well defined. At zero temperature, a system exists in its ground state; entropy is determined by how the ground state degenerates or it would be impossible to reduce any system to absolute zero in a finite number of operations. This law gives a reference point that is absolute for determining entropy. The entropy that is determined to be relative to this point is the absolute entropy.

Zeroth Law of Thermodynamics

The "Zeroth" law states that if two systems are in thermal equilibrium at the same time with a third system, then they are in thermal equilibrium with each other. If A and C are in thermal equilibrium with B, then A is in thermal equilibrium with B. In practicality, this says that all three are the same temperature. This forms a basis to compare temperatures. Any system with a number of microscopic parts, such as a gas, that is isolated from all forms of energy exchange and left alone for quite awhile will move toward a state of thermal equilibrium. A system in thermal equilibrium is characterized by a set of quantities that are macroscopic and are dependent on the system in question and characterizes its state. Two systems are in equilibrium, if when put in thermal contact with each other, their variables do not change.

Bridgman's Thermodynamic equations

Bridgman's Thermodynamic equations produce a large number of thermodynamic identities that involve a number of thermodynamic quantities. The variables are:

- U = Internal energy
- F = Hemholtz free energy
- G = Gibbs free energy
- H = Enthalpy
- N = Particle number
- P = Pressure
- p = Density
- S = Entropy
- T = Temperatures
- Cv = Heat capacity, constant volume
- Cp = Heat capacity, constant pressure

Many thermodynamic equations are expressed in terms of partial derivatives. The partial derivative of enthalpy with respect to temperature while holding pressure constant may be written as:

$$C_P = \frac{(\partial H)_P}{(\partial T)_P}$$

Rewriting of the partial derivative allows the use of many thermodynamic equations such as:

$$(\partial H)_P = C_P$$

Chemical potential energy and electromagnetic potential energy

Everything has molecules. Energy is required to make these molecules and hold them together. The energy stored in molecules is called chemical potential energy. An example is the energy stored in gasoline. Bonds are broken and reformed during combustion and new products are made. The energy stored in gasoline is released when it is burned, which is combustion. Gasoline is changed into byproducts during combustion such as water and carbon dioxide, and energy is released. An airplane motor will use the energy that is released to turn a propeller. A battery has chemical potential energy as well as electrical potential energy. When a flashlight is turned on, the electrical potential energy stored in the battery is converted into other forms of energy such as light. With an electrical appliance that is plugged in, electrical potential energy is maintained in a power plant's generator, a windmill or a hydroelectric dam.

Chemical reactions

The classic chemical reaction is a transfer of electrons resulting in a transformation of the substances involved in the reaction. The changes may be in composition or configuration of a compound or substance, and result in one or more products being generated which were not present in isolation before the reaction occurred. For instance, when oxygen reacts with methane (CH_4), water and carbon dioxide are the products; one set of substances (CH_4 + O) was transformed into a new set of substances (CO_2 + H_2O). Reactions are classified in many ways, some of which are as follows: as combination or synthesis, in which two or more compounds unite to form a more complex compound; decomposition, in which a compound is broken down into its constituent compounds or elements; and isomerization, in which compounds undergo structural changes without changing their atomic composition.

Some of the different types of chemical reactions are described below:

- Combination reactions: when pairs of reactants combine to produce a single substance. The reactions take place when it the energy is favorable to do so. For instance:
    - $C(s) + O_2(g) \rightarrow CO_2(g)$
    - $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
    - $CaO(s) + H_2O(1) \rightarrow Ca(OH)_2(s)$
- Decomposition: refers to the reaction involving those molecules which are stable at room temperature and decompose when heated:
    - $2KC1O_3(s) \rightarrow 2KC1(s) + 3O_3(g)$
    - $PbCO_3(s) \rightarrow 2PbO(s) + CO_2(g)$
- Single substitution: when a reaction involves an element that displaces another in a compound such as when a copper strip displaces silver atoms and produces copper nitrate and precipitating silver crystals of metal:
    - $Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$
- Double substitution: when a reaction looks as if it is exchanging parts of the reactants. An example:
    - $2KI(aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbI(s)$

Mechanisms

Chemical reactions normally occur when electrons are transferred from one atom or molecule to another. Reactions and reactivity depend on the octet rule, which describes the tendency of atoms to gain or lose electrons until their outer energy levels contain eight. Reactions always result in a change in composition or constitution of a compound. They depend on the presence of a reactant, or substance undergoing change, a reagent, or partner in the reaction less transformed than the reactant (such as a catalyst), and products, or the final result of the reaction. Reaction conditions, or environmental factors, are also important components in reactions. These include conditions such as temperature, pressure, concentration, whether the reaction occurs in solution, the type of solution, and presence or absence of catalysts. Reactions are described with the following equation:
Reagent(s)
Reactant(s)Product(s)
Reaction
conditions

Reactions

<u>Classes</u>
Organic chemistry and biochemistry are the study of large numbers of compounds involved in complex reactions. These classes of reactions consist of broad categories containing many subsets.

- Addition reactions are the combination of two reactants into a single, more complex product. Atoms are added to adjacent atoms forming a new compound.
- Elimination reactions occur when a single compound is split into two smaller compounds, one of which is usually a simple molecule such as H_2O or HCl. Atoms are eliminated from the original compound and usually bond to adjacent C atoms in the reactant(s).
- In Hydrolysis reactions a large molecule reacts with water and is split into two smaller molecules. These are the reactions which change starch into glucose for metabolism.
- Hydrogenation reactions occur when gaseous H combines with hydrocarbons and produces a saturated compound—a compound in which the C atoms are bonded to the maximum possible number of H atoms.

Alkanes and cycloalkanes

Alkanes are a class of saturated hydrocarbons which have formed themselves into long chains. Cycloalkanes are saturated hydrocarbons in the shape of rings. A Greek prefix before the name indicates the number of C atoms in the alkane; e.g. heptane contains seven C atoms. Because alkanes and cycloalkanes are highly stable, they are not very reactive because electrons are held in small orbits around the C and H atoms, making them less available for reactions. They can react under extreme conditions with substantial reaction energy to get them started. These reactions are generally classified into two groups: combustion, which is the often explosive reaction of alkanes with O, and halogenation, in which a halogen atom (from Group 17 on the periodic table) takes the place of an H atom in the presence of light.

Alkanes are less dense than water, which explains why gasoline, which is largely a collection of alkanes, floats on water. The molecules are nonpolar and are not soluble in water, which is polar. Their nonpolarity makes alkanes solvents of other nonpolar organic substances such as fats and oils following the “like dissolves like” rule of solubility. The boiling point of alkanes increases with increasing complexity and molecular weight. For example, methane, a simple, unbranched alkane boils at –164° C, while decane, a complex branched alkane boils at 174.1° C. Cycloalkanes are similar in physical properties in the predictability of density and boiling points. The higher the molecular weight of alkanes and cycloalkanes, the higher their respective density and boiling point.

Homologous series: A homologous series is a group of organic compounds sharing the same general formula, the same functional groups (and thus similar physical properties), and which differ predictably and incrementally as their relative size and atomic mass change. The most dramatic examples of a homologous series are the alkanes (methane, ethane, propane, butane, pentane, and so on). Each successive alkane differs from its predecessor in the presence of one additional CH_2 group, so the alkane's size and molecular mass increase accordingly. Alkenes and alkynes also appear in homologous series and exhibit the same periodic trends based on predictable changes at each successive step. It is important to remember that homologous series, while having exactly the same composition and proportions of constituent chemicals, may appear in many different isomers and thus have very different properties.

Alkenes

Alkenes are hydrocarbons structurally related to alkanes. They are characterized by C=C double bonds, making them unsaturated, and are very common in nature. Reactions associated with them typically involve the addition of functional groups or elements to the –ene parent hydrocarbon. One important reaction of this type is called hydrogenation because hydrogen gas combines with the alkene. The double C=C bonds are broken resulting in the formation of a saturated alkane. This process is also called alkene reduction because the double bonds are reduced to single bonds. This reaction has a high activation energy, and therefore is a slow reaction. Several catalysts have been synthesized which will speed up these reactions. In biological reactions such as those occurring in cells, hydrogenation or reduction of alkenes occurs with the aid of complex enzymes which serve as catalysts and sources of H.

Alcohols

Alcohols are a functional group containing oxygen and hydrogen (—OH, also called a hydroxyl group) bonded to the parent hydrocarbon with a single bond. An alcohol has multiple sites of reaction and different reaction types in which specific bonds are broken. In acid-base reactions, the O—H bond is broken. Depending on conditions, alcohols can act as acids or bases because they can be proton donors or acceptors. The C—O bond breaks in substitution reactions in which the hydroxyl group is replaced by something else. The C—O bond and C—H bond of an adjacent C atom are broken in elimination reactions in which a single molecule splits to form two products. The O—H and C—H bonds of the hydroxyl bearing atom are broken in oxidation reactions in which oxygen combines with C in place of H atoms.

An alkoxy (or alkoxyl) group is any group and an oxygen atom attached to an ether. It can react with hydrogen to form an alcohol.

Dehydration is a very important reaction involving alcohols in which H and O atoms are removed as H2O. It is an elimination reaction because two or more covalent bonds are broken in a single alcohol molecule and a C=C bond is formed in the product, forming an alkene. If two or more alcohol molecules are involved in the dehydration reaction, an ether is the product.

Ethanol (CH3CH2OH) is the alcohol in alcoholic beverages and is a product of fermentation. Other substances are sometimes added to ethanol to make it unfit for drinking; such altered ethanol is called denatured alcohol.

Primary, secondary, and tertiary alcohols are those in which the alkoxy group (–OH) is bonded to the primary, secondary, or tertiary C atoms, respectively.

Aldehydes and ketones

Aldehydes are carbonyl groups (C=O) to which at least one hydrogen atom is attached. Formaldehyde is the simplest known aldehyde, with two H atoms attached to the C (COH2). Ketones occur when a carbonyl C atom is bonded to two other C atoms, which are often parts of long hydrocarbon structures such as alkanes or cycloalkanes. Aldehydes and ketones are common classes of chemicals, especially in non-biological organic molecules. In common names, suffixes identify the compounds as aldehydes (-al) or ketones (-one), such as ionone, an aromatic compound found in irises and perfumes. Common reactions include reduction reactions, in which both aldehydes and ketones can be reduced to alcohols by breaking the C=O double bonds and adding H. By contrast, aldehydes engage easily in oxidation reactions, as in the addition of O- , to form carboxylic acid; ketones are not easily oxidized.

Acetal is a product of an aldehyde's reaction with an alcohol and is an important compound in carbohydrate chemistry. Hemiacetals are unstable intermediates on their way to becoming acetals; their transformation is completed in the presence of an acid catalyst.

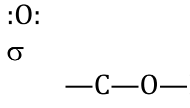
Benedict's reagent is an important component of experiments involving aldehydes. It is a solution with mild oxidizing properties containing Cu^{2+} ions and produces a red precipitate in the presence of aldehydes.

Ketals are products of a ketone's reaction with alcohol. Hemiketals are the intermediates in such reactions and are very unstable.

Tollen's reagent is used for similar purposes but contains silver ions instead of copper. In the presence of aldehydes, the silver in solution precipitates to form solid silver.

Carboxylic acids

Carboxylic acids are compounds formed when a carboxyl group (CO_2H) bonds to a hydrocarbon unit such as an alkane or alkene. Formic, acetic, and stearic acids are common carboxylic acids. In decarboxylation of these compounds, a substitution reaction, the carboxyl group is replaced by a hydrogen atom. These reactions are important in biological processes. These acids may also be converted to anhydrides (compounds without water) by reactions, usually at high temperature, which break the bonds in the carboxyl group and produce H_2O. Cleaving the C=O bond and removing the OH group from the carboxyl results in an acyl with double or triple bonds between the C and remaining O. These groups react quickly with halides to form compounds such as acyl chloride (OCCl).



$$-\overset{\overset{:O:}{\sigma}}{C}-O-H$$

Carboxyl

Benzenes

Benzene (C_6H_6) is an important cyclic organic compound. The electrons of its C atoms are delocalized resulting in combinations of single and double bonds, a structure which gives benzenes unusual stability. Compounds with strong, usually pleasant odors are often attached to benzene parents, so early chemists named these compounds aromatic compounds. Many benzene compounds have no odor, however. Common reactions include the following:

Aromatic electrophilic substitution: the most important class of reactions associated with benzene. In these reactions one H atom of the benzene is replaced by an electrophile which can be an atom or, more often, a molecule or functional group. This class contains subclasses of reactions such as halogenation (benzene + a halogen), nitration (benzene + a nitrate), sulfonation (benzene + a sulfate), alkylation, and acylation.

Amines and amides

Amines ($R—NH_2$) are a functional group containing nitrogen which form compounds vital to biological processes. They form when a hydrogen from an ammonia molecule (NH3) is replaced by an organic compound. They can be primary, secondary, or tertiary amines, depending on how many atoms are bonded to the N atom. Amines react with acyl chlorides (Cl—C=O) and acid anhydrides in nucleophilic substitution, producing amides. They neutralize carboxylic acid in an acid-base reaction to produce ammonium carboxylate. Tertiary amines react with strong halogen acids to form ammonium salts. Amides (O=C—N) are another functional group; the most important of their reactions is hydrolysis in which the addition of water severs the C—N bond to create amine and carboxylic acid.

Esters

Esters (O=C—O) are a functional group often derived from carboxylic acids from which a H atom has been removed. This process is usually a condensation reaction in which two hydrogen-containing groups or compounds react, joining their H and releasing it as water. Such condensation reactions that produce an ester are called esterification reactions. The most important class of reactions associated with esters is hydrolysis, a reaction in which the ester is broken down by water to yield an alcohol and a carboxylic acid or salt. Hydrolysis of esters may be catalyzed by an acid or a base; if the catalyst is a base, the process if called saponification. Esters may also form amides through a reaction with primary or secondary amines at high temperatures.

Ethers

Ethers are a functional group containing an O atom bonded to two alkyl groups. “Ethers” is also the name of the class of chemicals containing an ether group. They are somewhat soluble in water because the O bonds readily with H and they have a relatively low boiling point. Ethers can be derived from the dehydration of alcohols through heat and a hydrogen-rich catalyst. Nucleophilic displacement of alkyl halides also yields ethers. Ethers are not very reactive in the pure state, but their substituents groups can be highly combustible, as with diethyl ether. Hydrolysis breaks ethers down, but only under drastic conditions such as boiling in the presence of strong halogen acids. Ethers form explosive peroxides in reactions with adjacent CH groups.

Acyl halides

An acyl halide is a compound derived from either a hydrogen acid or a carboxylic acid in which a hydroxyl group (–OH) has been replaced by a halogen. A double bond forms between C and O with two single bonds joining the C with the halogen atom and a root molecule. Acyl halides are synthetic molecules and are most often used as intermediates in reactions to form other compounds. In hydrolysis reactions, they form carboxylic acid and a hydrogen halide. They are useful in forming esters through a substitution reaction with alcohol and amides by reacting with amines. They also react with benzene (aromatic) rings to form an aromatic ketone using a strong acid catalyst. Acyl halides are quite reactive and somewhat toxic.

Enantioselectivity, regioselectivity, and diastereoselectivity

Reactions in which a certain type of product is preferred, though various isomers are equally possible, are called selective reactions. A prefix before selective identifies the type of isomer predominately formed in such selective reactions. Enantioselectivity refers to the tendency of a reaction to form a specific enantiomer (nonsuperimposable stereoisomer). Regioselectivity is the same principle applied to “regions” of the molecule; some reactions produce more compounds with a specific orientation of appended groups than isomers with different orientations. Diastereomers are isomers which are not mirror images of each other and can exhibit different properties and reactivity. Diastereoselectivity describes reactions in which one stereoisomer is preferred over another as product; the reaction produces more of the favored variety.

Metabolic pathways

Metabolism refers to an organism’s ability, through chemical reactions, to produce and/or consume energy. Metabolic pathways are complex series of reactions in living cells which enable the biological processes. The pathways are highly complex, but well ordered and predictable, consisting of a series of chemical reactions designed either for catabolism—the breakdown of biomolecules, usually to produce energy, or anabolism—the synthesis of biomolecules. Photosynthesis is a widely known metabolic pathway which allows green plants to use energy from sunlight, water, and nutrients to produce the energy they need to live and grow. The processes commonly referred to as digestion, by which the human body extracts energy from food, is a series of enormous metabolic pathways. Because the chemical reactions involved are often not species-specific, understanding a metabolic pathway on a molecular and chemical level applies to other organisms that use the same process.

Factors influencing reactivity

In addition to describing the tendency of a substance to react, reactivity also refers to the speed with which reactions take place. Several factors influence this aspect of reactivity, including the purity of a compound or the presence of contaminants, the surface area of a substance, and the crystal structure of a substance that is in a solid state. An element's electron configuration is also an important factor because atoms "want" to reach their most stable possible configuration, so they react with other atoms in order to achieve this. Atoms with single, unpaired electrons, such as H, react quickly with other elements to achieve a more stable configuration - such as the formation of diatomic hydrogen, in which the orbital is filled to its capacity.

Photosynthesis

Photosynthesis is the process, in plants, of capturing light energy and using it to convert carbon dioxide and water into sugars. Its generalized formula is:

$$6\ CO2 + 6\ H2O + \text{light} \rightarrow C6H12O6 + 6O2$$
$$(keq=10\text{-}500)$$

Photosynthesis is a light-driven, oxidation-reduction reaction. This is known as a "redox reaction." Redox reactions involve the stable transfer of electrons between atoms. The primary molecule that is responsible for catching light energy for the photosynthesis process is chlorophyll. Chlorophyll makes a transition to a higher state of energy when it absorbs a photo of light:

$$\text{chl} + \text{a photon} \rightarrow \text{chl}^*$$

The energy can be dissipated as fluorescence, heat, resonance transfer or redox reaction.

Laboratory-induced chemical synthesis

Chemical synthesis is the intended use of chemical reactions to produce a product or products. This may take place by using manipulations that are physical and chemical. It usually involves one or more reaction. In laboratory usage, this implies the process can be reproduced, will be reliable and can be established to take place in other laboratories. The synthesis starts by selecting a compound that is a reactant. Various types of reactions may be applied in order to synthesize a product or a product that is intermediate. Reaction yield refers to the amount of product a chemical synthesis makes. Yields are usually expressed in mass as grams or as a percentage of a theoretical quantity of a product that might be produced. Types of synthesis include total synthesis, organic synthesis and telescopic synthesis.

Chemical synthesis as a chemical reaction

Chemical synthesis can also define a somewhat narrow and restricted type of specific chemical reaction that is a direct combination reaction in which two or more reactants combine and form a single product. The form for this reaction is:
A + B -> AB where A and B are elements or compounds, and AB is also a compound containing A and B. An example is the formation of table salt:

$$2Na + Cl2 \rightarrow 2\ NaCl$$

There are four special synthesis rules:

- Metal-oxide + H2O -> metal (OH)
- Nonmetal oxide + H2O -> oxi-acid
- Metal chloride + O2 -> metal chloride
- Metal-oxide + CO2 -> metal (CO3)

Energy

Energy is required for chemical reactions to occur and is released during reactions. One way to think of reactions is the breaking and formation of chemical bonds. Because chemical bonds have measurable strength, they tend to stay together unless acted upon by an external force, or energy; this energy is the reaction's activation energy. When bonds are broken or formed, energy stored in the bonds or substances is released. Energy on both sides of reactions occurs in various forms such as kinetic energy, heat, and light, and its measurement is an important part of understanding and balancing reactions and equations. Measuring the energy of reactants compared to that of products yields a difference in energy, or enthalpy, symbolized in equations by ΔH°, which may be thought of as the "heat" of a reaction (though energy does not always take the form of heat).

Transition state, intermediate, concerted reaction, and multi-step reactions

The transition state is the point of highest energy during a chemical reaction. Reactions exhibit differing energies—or enthalpies—as a reaction progresses; these differences are plotted as a curve in "reaction coordinate diagrams" and are one way of tracing the reaction rate. The transition state is the highest point on the plotted curve. An intermediate is a state of matter produced during a reaction that is neither reactant nor product, but something intermediary. They are typically highly unstable and exist for only a short time before becoming product. Concerted reactions are reactions in which bonds break and form simultaneously with no time interval or intermediates. Multi-step reactions require several steps to complete and often involve intermediates and enzymes, and may also depend on the polarity of the solution.

Catalysis

Catalysis occurs when a catalyst is added to a reaction to increase its rate and efficiency. Catalysts change the rate of a reaction without being changed *by* the reaction and are important elements of chemical and biological processes. They function by decreasing the activation energy required for a reaction to occur. Homogeneous catalysis occurs when the catalyst is in the same state of matter as the reactant(s) and product(s). Heterogeneous catalysis occurs when the catalyst is in a different state. For example, hydrogen peroxide (H_2O_2) breaks down naturally to produce water and oxygen; this is normally a very slow reaction. Liquid hydrogen bromide can be added to the H_2O_2 to speed up the reaction considerably; because all compounds in question are liquid, this is an example of homogeneous catalysis. Solid manganese dioxide crystals (MnO_2) can also be added to the H_2O_2 as a catalyst to demonstrate heterogeneous catalysis.

Reaction enthalpy

Enthalpy is the expression of the chemical energy in elements and/or compounds, expressed in equations by H. Systems with high energy being said to have high enthalpy. This is measured by comparing the energy of a system to another whose enthalpy is known to be zero. The difference between the two gives the enthalpy of the system in which H is unknown, much as altitude is determined by using sea level as the constant 0 value. Because enthalpy (or the difference in enthalpy between products and reactants) is measured in final and initial states, the value will be the same regardless of reaction pathway. Enthalpy and its related measurements assume constant pressure. This is important to remember because all matter expends energy to "push back" against the atmosphere; enthalpy does not include this energy.

Electrophiles, nucleophiles, pericyclic reactions, and electrocyclic reactions

An electrophile is the acceptor of an electron pair in chemical reactions. A cation "seeks" additional electrons to balance its negative charge, so it is an electrophile. Nucleophiles are donors of electron pairs in reactions. Anions are nucleophilic because they "seek" a positively charged region near a nucleus to balance their negative charges. Pericyclic reactions are reactions whose defining characteristic is the reorganization of π and σ bonds through cyclic transition states. There are no intermediates in these reactions because electrons shift during the transition state to transform the reactant immediately to product. Or, the molecular orbitals of the reactant are transformed into the molecular orbitals of the product. Pericyclic reactions require activation energy, though not necessarily a great deal. Electrocyclic reactions are a class of exothermic pericyclic reactions in which a cyclic reactant is transformed into another cyclic product with one less double bond than the reactant.

Electrophilic and nucleophilic addition reactions

Addition reactions are those in which two reactants combine to yield a single product; they are characterized by the cleavage of double or triple bonds in the central molecule. Electrophilic addition reactions are those that depend upon the presence of an electrophile, often a proton. For example, in the addition reaction of HBr with an alkene, the alkene takes a proton from the HBr molecule (an acid) to form an intermediate C atom with a positive charge. That intermediate then combines with the Br- ion to form the final product—an alkane. Addition reactions are nucleophilic when a negative ion donates an electron pair at an electrophilic site, severing the double or triple bond to create a single product. In the reaction of HCl with an alkene, the C=C bond reacts with H to form H—C—H, a C^+ ion, and a Cl- ion, which donates an electron pair to C^+, neutralizing both ions.

Electrochemistry

Electrochemistry describes the movement and generation of electrical charges or forces in chemical reactions. Electrical current can be used as activation energy as in electrolysis—the process of separating compounds into elements using electricity; for instance, water can be split into its constituent gases by passing an electrical current through it. Electricity can also be the desired product of reactions such as the acid reactions occurring inside batteries which gives them an electrical charge. Electricity can be envisioned as the movement of large numbers of electrons, all of which carry a negative electrostatic charge. The reactions in batteries break bonds and free electrons from their atoms so they can flow through conductors to the load, or destination of the electricity.

Voltaic cells

Voltaic cells, also known as galvanic cells, are important to electrochemistry because they were the first experiments to describe the flow of electricity produced by chemical reactions. They are the precursors of modern batteries. The cells involve two chemical mixtures and an oxidizing metal. In one mixture, a strip of solid zinc is placed in a zinc sulfate solution; and in the other, a strip of copper is placed in a solution of copper sulfate. An electrical current is generated by an oxidation reaction on the Zn side in which electrons are freed. A reduction reaction occurs on the Cu side, causing the solid Cu to gain electrons. A wire connects the two strips and allows electrons to move from the Zn side to the Cu side, generating an electrical current.

Conduction

Conduction in metals is due to electrons in conduction bands. Conduction in insulators is due to electrons in conduction bands and valence band holes. Holes are vacant states in valence bands that are removed when created. A definition for the difference between insulators and semiconductors is electrons at ordinary temperatures can reach the conduction band. The probability that an electron reaches the conduction band is about $\exp(-Eg/2kT)$ where Eg is the band gap and kT has a usual meaning. In addition to having relatively small Eg, semiconductors have a covalent bond and insulators usually have a partial ionic bond.

Atomic spectra dynamics

Electrons in atoms have sets of allowed energy levels. They gain or lose energy by traveling between levels and absorbing emitting photons. Zero energy is achieved in an electron that is resting at infinity. Each atom has a different set of negative energies allowed. Atoms can drop from an allowed energy E1 to a lower E2 by the emission of a photon with energy.

$$hf = E1 - E2$$

Frequencies that correspond to the differences between allowed energies may be emitted. An atom can move to a higher allowed energy by absorbing an energized photon.

$$hf = E2 - E1$$

When atoms are close to each other, the energy level splits as electrons interact. In a solid, a continuous frequency range can be emitted or absorbed. Hot and dense materials emit continuous spectra that hold all frequencies. An example would be a light bulb filament emitting light at all frequencies.

Balmer and Rydberg equations

The Balmer formula is one of a set of six individual series that describes spectral line emissions of the hydrogen atom. Balmer saw a single number related to every hydrogen spectrum line in the visible light region. This number is 364.56 nm. The formula can be used for finding the wavelength of the absorption/emission lines and was originally written:
Where λ is the wavelength, h is a constant with a value of 3.6456×10-7 m or 364.56 nm; n is equal to 2; and m is an integer that is > n.

$$\lambda = B\left(\frac{m^2}{m^2 - n^2}\right) = B\left(\frac{m^2}{m^2 - 2^2}\right)$$

Rydberg made the formula for all transitions of hydrogen:

$$\frac{1}{\lambda} = R_{\mathrm{H}}\left(\frac{1}{2^2} - \frac{1}{n^2}\right), n = 3, 4, 5, \ldots$$

Where λ is the wavelength of the absorbed/emitted light and RH is the Rydberg constant with the constant for an infinitely heavy nucleus is 3.6456 x 10-7 m or 364.56 nm.

H-alpha

H-alpha, also written, Hα, is an emission line that is created by hydrogen. Electrons exist in quantized energy levels that surround the atom's nucleus according to the Bohr model. The energy levels are described by the principal quantum number, n = 1, 2, 3 ... Electrons may only exist in these states and may only transition between states. The set of transitions from n ≥ 3 to n = 2 are the Balmer series; n = 3 to n = 2 is known as Balmer-alpha or H-alpha; n = 2 is H-beta; n =5 to n = 2 is H-gamma and so on. The Lyman series is named n = 2 to n =1, known as Lyman-alpha; n =3 to n = 1 is Lyman-beta and so forth. H-alpha's wavelength of 656.3 nanometers is visible in the red portion of the electromagnetic spectrum.

Electron spin resonance

Electron spin resonance, ESR, is a spectroscopic technique used to detect species with unpaired electrons. Generally it means that it has to be a free radical of an organic molecule or a transition metal ion if it's an inorganic complex. Also known as electron paramagnetic resonance, EPR, the technique is rarely used in nuclear magnetic resonance (NMR) because most stable molecules have a closed-shell configuration without an appropriate unpaired spin. Weak magnetic fields and higher frequencies are used in comparison with NMR because of the difference in mass between the nuclei and electrons. ESR is used to identify and quantify radicals, or molecules with unpaired electrons. In chemistry it is used for identifying reaction pathways and is also used in medicine for biological spin probes.

Law of Conservation of Mass

The Law of Conservation of Mass in a chemical reaction is commonly stated as follows:

In a chemical reaction, matter is neither
created nor destroyed.

What this means is that there will always be the same total mass of material after a reaction as before. This allows for predicting how molecules will combine by balanced equations in which the number of each type of atom is the same on either side of the equation. For example, two hydrogen molecules combine with one oxygen molecule to form water. This is a balanced chemical equation because the number of each type of atom is same on both sides of the arrow. It has to balance because the reaction obeys the Law of Conservation of Mass.

Law of Conservation of Energy

The Law of Conservation of Energy is simply a restatement of the First Law of Thermodynamics. It maintains that the total energy that flows into a system has to be equal the total energy that flows out of the system, plus any change in the energy contained within the system. A more common formulation is that energy can be neither created nor destroyed, it can only be converted from one form to another.

Kinetics

Reaction kinetics

“Reaction kinetics” is the study of the speed at which reactions occur and the sequence of bond formation and bond cleavage occurring in reactions. This concept is important because understanding the speed and sequence of reactions for a few samples of a given class of chemicals enables predictions about the reaction mechanisms of the entire class and compounds in which that class is found. “Kinetic” refers to motion and the energy inherent in moving bodies. Reactions occur when particles move about within a substance and react with one another, often using the kinetic energy as activation energy. The higher the speed of particles in motion, the higher the kinetic energy; this tendency explains why reactions often occur more quickly at higher temperatures.

Rate laws, first order reactions, second order reactions, and zero order reactions

Rate laws describe the relationship between the concentration of reactants and reaction rate. The relative speed of a reaction may increase with changes in reactant amounts; it may also decrease or show no impact. Rate laws track the effects of changes in reactant concentration on reaction rates. First order reactions are those reactions in which only a single reactant has an impact on reaction rate. The rate increases or decreases in direct proportion to the concentration of the reactant: doubling the concentration of reactant doubles the rate. Second order reactions are those in which both reactants impact the rate of reaction, and the relationship is not proportional, but exponential: doubling a reactant concentration raises the rate by four. Zero order reactions are those in which changes in reactant concentration have no effect on reaction rate.

Reaction rates

Many factors influence the rate of chemical reactions, but the three most common factors are temperature, subdivision, and concentration:

- The higher the temperature, the faster the rate of reaction because reactant particles are moving more quickly and collide with other particles more often. Furthermore, particles are moving with greater energy at higher temperatures, increasing the chances that collisions result in reactions.
- Subdivision of reactants increases the rate of reaction because greater surface area means more particles are available to react. A solid block of baking soda will react with vinegar more slowly than powdered baking soda because less surface area is available, meaning fewer collisions of particles.
- Concentration of reactants influences the rate of reaction because greater concentration means more particles are present

to collide and react with other particles.

Reaction rates can only be determined with certainty through experiments. Estimates can be made based on known properties of a compound, laws of equilibrium, and reaction conditions; but precise data only come from experiments. Some data considered in reaction kinetics are the amounts of reactant(s) and product(s), specifically the rate at which a reactant is being consumed (or changed via the reaction) and the rate at which product emerges. These amounts and the time period can then be plotted on graphs to provide a curve measuring the speed of the reaction. Measuring the concentration of a substance over time is another way to track a reaction's rate. For instance, the changes in volume of a product measured at ten minute intervals give an accurate average speed of reaction, assuming initial quantities are known.

<u>Factors influencing rate</u>
Chemical reactions can vary in how quickly they occur. These rates depend on many factors, including the following:

- The number of collisions of particles. Chemical reactions occur when electrons are transferred from one particle to another. Therefore, the greater the number of collisions between particles, the greater the transfer, or reactivity. Several factors determine the number of collisions including temperature (particles move faster and collide more at higher temperatures), concentration (more particles = more collisions), and the presence of other substances which facilitate collisions.
- The orientation of particles at the time of collision. Not all sites on a particle are reactive; if a collision occurs at a non-reactive site, the reaction is impeded.
- The energy with which the collisions occur. Just as damage is greater in a car crash at higher speeds, greater energy is present when particles collide at higher velocities, increasing the reactivity.

Conservation of momentum

Momentum seems to constantly appear conserved. Without outside forces, a system will have an omnipresent total momentum or a property implied by Newton's law of inertia, which was his first law of motion. Newton's law of reciprocal actions states that forces acting in between systems are equal, which is equal to the conservation of momentum. Momentum also has special properties in which inside a closed system it always is conserved although it is not conserved in inelastic collisions. Elastic collisions, such as two pool balls hitting each other, conserve kinetic energy. Inelastic collisions do not conserve kinetic energy. An example of this is when two objects strike each other and stick together after the collision.

Cosmic rays

Cosmic rays are radiation that are made up of energetic particles that originate beyond the Earth but impinge on the Earth's atmosphere. Cosmic rays are mainly made of ionized nuclei, the majority of which are protons. Kinetic energy from cosmic ray particles span over 14 orders of magnitude with the flux of cosmic rays on the Earth's surface falling at about the inverse cube of the energy. A wide variety of particle energy reflects various sources. Cosmic rays start from processes on the Sun and extend to the farthest portions of the universe that is visible. They may have energies up to 1020 eV. Victor Hess, using electrometers to measure ion rates inside a sealed container, flew to 5,300 meters in a balloon in 1912 and found a four-fold increase in ionization over that at ground level. He explained it by assuming that great penetrating radiation enters the atmosphere.

Half-life of a radioactive isotope

The half-life of a radioactive isotope is the time required for one half of the original atoms in a sample to decay. It is the principal characteristic used to distinguish one radioactive substance from another and can vary greatly from nucleus to nucleus. A quantity is said to subject to exponential decay if a decrease occurs at a rate that is proportional to its value. All radioactive decays obey first-order kinetics. Those quantities subject to exponential decay are known by the symbol N. If the quantity is denoted by N, the value of N at a particular time t is expressed by:

$$\frac{dN}{dt} = -\lambda N.$$

The solution is:

$$N = Ce^{-\lambda t}.$$

This is the form of the equation most commonly used to describe exponential decay. The constant of integration C is often written N0 since it stands for the original quantity.

Alpha decay

Alpha decay is a type of radioactive decay in which an alpha particle is ejected from an atomic nucleus through the electromagnetic force and results in a nucleus with a mass number 4 less and an atomic number 2 less. An example:

$$^{238}\mathrm{U} \rightarrow {}^{234}\mathrm{Th} + \alpha.$$

The alpha particle is a helium nucleus, both atomic number and mass number are conserved. Alpha decay can be looked at as nuclear fission, in which the parent nucleus splits into a pair of daughter nuclei. Alpha decay is governed by the strong nuclear force, or the force between two or more nucleons. Alpha particles have speeds of 15,000 km/s when a typical kinetic energy of 5 Me V.

Neutrinoless double beta decay

Double beta decay is also called two neutrino double beta decay since two neutrinos or anti-neutrinos are emitted. If the neutrino is a Majorina particle in which the anti-neutrino and the neutrino are one and the same particle, then a neutrinoless double beta decay could possibly take place. When neutrinoless double beta decay occurs, the emitted neutrino is quickly absorbed as its anti-particle by another nucleon of the nucleus. So, the total kinetic energy of the two electrons are exactly the difference in binding energy that is between the initial and final nuclei states. Experiments to look for neutrinoless double beta decay have taken place or have been proposed. Such a discovery would indicate that neutrinos are Majorana particles which allow a neutrino mass calculation.

Classes of nuclear reactors

Nuclear chain reactions are started, controlled and sustained at a steady rate in a nuclear reactor. This is opposed to a nuclear explosion in which chain reactions occur in a split second. Reactors have many uses, the major one being electric power generation and sometimes for producing plutonium to be used in nuclear weapons. All commercial reactors in the U.S. operate on nuclear fission. Reactors are generally divided into two classes: thermal reactors using slow or thermal neutrons and fast reactors that use neutrons to sustain the fission chain reaction and are characterized by a lack of material for moderating. Thermal reactors have moderating materials which are meant to slow the neutrons until they approach the average kinetic energy of surrounding particles, in other words until they are thermalized. They need highly enriched, sometimes weapons grade, fuel or plutonium so a reduction can occur in the amount of U-238 that would otherwise capture fast neutrons.

Internal energy, thermal energy, and heat

A distinction is made in physics between heat, internal energy and thermal energy. Internal energy is all of the energy that belongs to a stationary system. This includes thermal energy, nuclear energy and chemical energy. The motion or interaction between atoms comprising the system is internal energy. If an outside or external object is involved in the interaction then it is not included in the internal energy. Thermal energy is the portion of internal energy that is changed when the phase of a substance changes or the temperature changes. This is due to the kinetic energies, or motion energies that are associated with the atoms' random motions in a system. Heat is simply the measure of thermal energy transfer, representing the change in thermal energy.

Kinetic gas theory

Kinetic gas theory maintains that gas pressure is the result of collisions between molecules and the walls of the container in which they exist. The total pressure depends upon the frequency of collision per unit area and the force with which the molecules strike the wall. The following assumptions are central to the theory:

- The gas molecules neither attract nor repel one another.

- The distance between the molecules is considered much greater than the size of the molecules. In other words, the molecules can be considered to be points possessing mass but with negligible volume.
- Gas molecules are in a state of constant, random motion. Collisions are perfectly elastic and although energy is transferred from one molecule to another because of the collision, the total energy of the system remains constant.
- The average kinetic energy of the molecules is proportional to the temperature in Kelvins. Two different gases will have the same kinetic energy provided they are at the same temperature.

Relative humidity

Relative humidity is the percent of saturation humidity that is calculated generally in relation to saturated vapor density.

Relative Humidity = actual vapor density/saturation vapor density x 100 percent

Evaporation in a closed container will proceed until as many molecules return to the liquid as escape. At this point, the vapor is said to be saturated. The pressure of that vapor, which is usually expressed in mmHg, is the saturation vapor pressure. Molecular kinetic energy is greater at higher temperatures, so more molecules may escape to the surface. The saturated vapor pressure corresponds as higher as well. If the liquid is open then vapor pressure is viewed as a partial pressure along with the other constituents of the air. The boiling point is the temperature at which the vapor pressure equals the atmospheric pressure.

Kinetic molecular theory

The kinetic molecular theory is based on these postulates:

- Gases are composed of a large number of particles. They behave like hard, spherical objects that are constantly in a state of random motion.
- These particles move in a straight line until they have a collision with another particle or the walls of the container.
- These particles are much smaller than the distance that is between the particles. Most of the volume of the gas is empty space.
- There is no force of attraction between the particles and the walls of the container or between gas particles.
- Collisions between gas particle or with the walls of a container are elastic. A gas particle loses no energy when it collides with another particle or the walls of the container.
- The average kinetic energy of a gas particle collection depends on the gas' temperature and nothing else.

The assumptions of the kinetic molecular theory can be shown in an experiment with a glass plate surrounded by walls that are mounted on top of three vibrating motors. About a handful of ball bearings are placed on top of the glass. This represents the gas particles. When the motors are started, the glass plate vibrates. This makes the ball bearings move in a random, constant way. Each ball moves in a straight line until colliding with another ball or with the walls. The collisions are frequent but the average distance between the ball bearings is much larger than the ball's diameter. There is no force of attraction between the individual ball bearings and the wall. Any object in motion has a kinetic energy that is defined as half the product of its mass times its velocity squared:
KE = 1/2 mv2

Amontons' Law

Amontons’ law is another postulate of the kinetic molecular theory. It states that the average kinetic energy of a gas particle is dependent upon the temperature of the gas. Consequently, the average kinetic energy of the gas particles increase as the gas gets warmer. Because the mass in these particles is constant, their kinetic energy will only increase if there is an increase in the average velocity of the particles. The faster that the particles move when they hit the wall, the greater the force that will be exerted on the wall. Since the force of each collision gets larger as the temperature rises, the gas pressure must also increase

Gay-Lussac's Law

Gay Lussac's discovered two laws. One states that between combining volumes of gases and the product, the ratio can be expressed in small whole numbers. Avogadro used this data to form his hypothesis. This paved the way for modern gas stoichiometry, or the quantitative relationship between products and reactants within a chemical reaction. Another Gay-Lussac Law states that the pressure of a fixed amount of gas at volumes that are fixed is directly proportional to its temperature in kelvins. It is expressed:

$$\frac{P}{T} = k$$

P is the gas pressure; T is the gas temperature in kelvins; k is a constant. The law is true because temperature is a way to measure average kinetic energy of a substance. As a gases' kinetic energy increases, a collision occurs between its particles and the container walls more rapidly, causing an increased pressure.

Equilibrium

Exothermic, endothermic, activation energy, and reaction equilibrium

Exothermic reactions are chemical reactions in which energy is released or produced, such as in combustion. In endothermic reactions, external energy is required for the reaction to occur and is absorbed from the surroundings. Some reactions require energy to start the reaction, this energy is called activation energy. A match applied to tissue paper is an example of activation energy. When the same number of atoms is present on the reactant side of an equation as on the product side, the reaction is said to be in equilibrium. All atoms of a substance must be accounted for after it undergoes a chemical reaction. For instance, methane reacting with oxygen produces water and carbon dioxide, but all atoms in the original substances are still present, albeit in different combinations.

Dynamic equilibrium

Dynamic equilibrium is a form of chemical equilibrium as applied to solutions. Solutes and solvents continually move around in solution and some particles of the solute reattach and take on their original forms; they "drop out" of solution. This process is called crystallization and is the opposite of dissolution. When the rates of dissolution and crystallization are equal, a solution is said to have achieved dynamic equilibrium. All solutions will eventually reach dynamic equilibrium unless they are subject to external energy, such as an increase in temperature or agitation (such as stirring) of the solvent. For example, a cube of zinc dissolves in HCl at a given rate. If the solution is left alone it will reach equilibrium when the rate at which particles of Zn leaving the cube equals the rate at which particles crystallized from the solution rejoin the cube.

Evidence of a chemical reaction

A gas may form or disappear, as when the calcium in an eggshell dissolves in vinegar's acetic acid (CH_3COOH). The gas appears as bubbles and will dissipate from an open container. A solid may form or disappear. When a lump of coal burns, the lump changes its form to ash, essentially disappearing. A liquid forms or disappears, as when flammable liquid burns, producing gases. A color change occurs, as when iron reacts with oxygen, producing rust. Energy is released as heat and/or light, as in the sudden combustion of gunpowder. The temperature of a solution may drop when energy is absorbed in the form of heat from the surroundings. When vinegar reacts with $NaHCO_3$ (baking soda), the temperature of the solution drops because the reaction absorbs and metabolizes heat from the outside.

Le Châtelier's Principle

Le Châtelier's Principle states that a change in a system at equilibrium changes the concentrations of reactants or products and a new equilibrium results. For example, if more reactant is added to a system at equilibrium, a proportional change will be made to the product side of the reaction, resulting in a new equilibrium; the change in reactant is offset by the change in product to minimize the disturbance to the system. Similarly, if a product is removed from a chemical system in equilibrium, reactants create more product until equilibrium is restored. The volume of products and reactants has changed, but the constants and proportions will return to balance. The principle is useful to predict how changes to a system will influence products and reactants.

Chemical equilibrium

While reaction equilibrium is the law which states that the number of atoms present in reactants must equal the number of atoms present in the product(s), chemical equilibrium has to do with the rates of reactions. Many chemical reactions "recycle" their products into reactants, which then become products again. When the rate of production equals the rate at which products are broken into reactants, chemical equilibrium is reached. The state is represented by a double, bi-directional arrow: Ý. When both arrows are of equal length, the amount of product(s) and reactant(s), as well as the rate, are balanced. Greater quantities on either side of the equation may also be maintained by a system in equilibrium. If the system favors the product side, the top arrow is longer. A longer bottom arrow indicates greater volume of reactant(s) in a balanced system.

Equilibrium constants

The equilibrium constant is the data in equations which describes the interplay of reactants and products in a system at equilibrium. Once equilibrium has been reached, the concentration or reactants and products and the rate of reaction will remain constant if reaction conditions remain unchanged. The constant is represented as K in equations, and subscripts may be added to further describe the state and relation of compounds. K_c is used when concentrations of reactants and products are known, and K_p is used when the compounds involved are in the gaseous state. The subscript c represents concentration; when considering gases, pressure is a more useful measurement than concentration and is represented by the subscript p. Using the ideal gas law, K_c and K_p can be converted back and forth as necessary.

Fermi energy

Fermi energy (EF) is the energy in a system of fermions that do not interact, has the smallest possible ground state energy (or its lowest energy state), and increases when only one particle is added to the system. It is equal to the chemical potential of the system in its ground state at absolute zero. It may also be seen as the maximum energy a single fermion can have in this ground state. Fermions obey Fermi-Dirac statistics, which is the statistical distribution of fermions over energy states in thermal equilibrium. The ground state of non-interacting fermion systems is built by beginning with an empty system and adding particles one after another and consecutively filling up the lowest-energy quantum states that are unoccupied. When the desired number of particles is reached, the Fermi energy is that of the highest occupied or lowest unoccupied state.

Acids and Bases

Definitions of acids and bases

A definition of acids has historically been difficult to pin down. Several names and theories have been influential in expanding our understanding of acids, including the following: Arrhenius acids are named for Svante Arrhenius, who first described acids and bases as proton donors or hydroxide ion donors, respectively. His theory proposed that acids' unique properties come from the disassociation, or ionization, of H atoms when in aqueous solution. The Brønsted-Lowry theory overcomes some limitations of the Arrhenius definition by defining acids as proton donors and bases as proton acceptors.

The Lewis theory focuses on electron pairs as the defining components of acids and bases; it proposes that acids accept an electron pair while bases give them up. Each model has some limitations; some substances exhibit acidic or basic properties that are not explained by the most useful model for that substance. According to the Lewis theory of acids and bases, acids are substances whose atoms or molecules can accept an electron pair from another substance, forming a covalent bond. Lewis bases form covalent bonds by donating an electron pair. The properties of Lewis acids and bases therefore do not rely upon dissociation of ions, so do not have to include O or H like other acids. Any substance with an incomplete outer energy level (i.e., an incomplete octet) will tend to form covalent bonds by accepting electrons, and can thus be considered a Lewis acid. In this reaction, water acts as a Lewis acid by accepting an electron pair, while N is the Lewis base: CN^- + H_2O Ý HCN + OH-

Physical properties of acids

Acids are a unique class of compounds characterized by consistent properties. The most significant property of an acid is not readily observable and is what gives acids their unique behaviors: the ionization of H atoms, or their tendency to dissociate from their parent molecules and take on an electrical charge. Carboxylic acids are also characterized by ionization, but of the O atoms. Some other properties of acids are easy to observe without any experimental apparatus. These properties include the following:

- A sour taste
- Change the color of litmus paper to red
- Produce gaseous H_2 in reaction with some metals
- Produce salt precipitates in reaction with bases

Other properties, while no more complex, are less easily observed. For instance, most inorganic acids are easily soluble in water and have high boiling points.

Common acids

Some common acids are often used in basic experiments because they are readily available and relatively safe and inexpensive. These useful acids include the following:

- Acetic acid. Formally known as ethanoic acid or methanecarboxylic acid; it is the acid responsible for the characteristic taste and smell of vinegar, household vinegar is simply acetic acid in aqueous solution. Its structural formula is CH_3COOH; its empirical formula is $HC_2H_3O_2$.
- Hydrochloric acid, HCl, is hydrogen chloride gas in aqueous solution. It is the main ingredient in gastric acid and has many industrial uses. It is highly corrosive, though not volatile by itself, so should be handled with appropriate lab safety precautions. It is nearly odorless.
- Nitric acid, HNO_3, is also a corrosive, strong acid with a sharp odor. It is used in the industrial production of fertilizers and explosives.

Acid/base strength

The characteristic properties of acids and bases derive from the tendency of atoms to ionize by donating or accepting charged particles. The strength of an acid or base is a reflection of the degree to which its atoms ionize in solution. For example, if all of the atoms in an acid ionize, the acid is said to be strong. When only a few of the atoms ionize, the acid is weak. Acetic acid ($HC_2H_3O_2$) is a weak acid because only its O2 atoms ionize in solution. Another way to think of the strength of an acid or base is to consider its reactivity. Highly reactive acids and bases are strong because they tend to form and break bonds quickly and most of their atoms ionize in the process.

Relationship to molecular geometry

Recall that the strength of an acid measures the tendency of its H (or O for carboxylic acids) atoms to ionize. In order to form the H^+ cation, a bond must be broken and particles transferred. Therefore, the easier it is to break this bond, the more acidic the substance will be. Molecular structure, or geometry, affects the ease with which bonds are broken in important ways. For instance, the size of the nonmetallic atom to which the H is bonded impacts bond strength because electrons in atoms with larger radii are more loosely held in orbit around the nucleus, leaving them vulnerable to reactions. Furthermore, the larger the molecule in which the H occurs, the easier it is for H^+ to separate from the molecule.

Hydrogen Theory of Acids

Hydrogen plays a central role in the modern understanding of acids and their properties; the set of ideas involving H atoms in acids is sometimes referred to as the Hydrogen Theory of Acids. These ideas focus on the hydrogen atom and its ability to dissociate, or separate from the rest of a molecule and become a H^+ cation. Not all H atoms in a substance can dissociate, and those that do can do so to different degrees—accounting for the relative “strength” or “weakness” of acids.

An acid is said to be strong when all or most of its available hydrogen atoms dissociate and become H^+. Acids such as H_2SO_4 (sulfuric acid) and H_3PO_4 (phosphoric acid) containing more than one dissociable H atom are called polyprotic.

Acid/base properties of water

Depending upon the reaction in question, water can act as either an acid or a base, making it an amphoteric substance. Pure water in a closed system exists at an equilibrium within itself in which small portions of it self-ionize and then return to the neutral state. This self ionization equilibrium is the foundational reaction to all solution chemistry and is expressed: H_2O (l) Ý H + (*aq*) + OH^-(*aq*). The reactions below demonstrate this amphoterism. Water donates a proton in the first, making it an acid, and accepts a proton in the second, making it a base.
NH_3(*aq*) + H_2O(l) Ý NH_4^+(aq) + OH^-(*aq*)
HCl(*g*) + H_2O(l)→H_3O+(*aq*) + Cl^-(*aq*)

Litmus and phenolphthalein

Litmus and phenolphthalein are both acid-base indicators. Litmus is a compound derived from certain types of lichens; it is often infused onto small strips of paper—litmus paper—for use in laboratory experiments. Litmus turns blue in contact with bases and red in contact with acids. It can also display a wide range of other colors indicative of various compounds and pH levels. Phenolphthalein ($C_{20}H_{14}O_4$) is an organic compound which also changes colors based on pH. In contrast to litmus, however, phenolphthalein is colorless in acid solutions and turns pink in basic solutions, with the color change occurring at around pH 8.3. It is also used industrially to make dyes and as a laxative, though toxic in large amounts. Indicators work through simple chemical changes and change colors because the changing molecular geometry in the reaction means the molecules reflect different wavelengths of light.

Neutralization

Neutralization is a chemical reaction that is also known as water-forming reaction. It is when an acid and a base reacts and produces a salt and water. It is a combination of hydrogen ions H+ and hydroxide ions OH- or oxide ions O2- to form water molecules (H2O). Salts are also formed during this process. Neutralization produces heat, or exothermic. The following normally occurs:

- Acid + Base --> Salt + Water : $\Delta H = -C < 0$

To use an example, take the reaction between sodium hydroxide and hydrochloric acid:

- hydrochloric acid + sodium hydroxide -> sodium chloride + water
- HCl(aq) + NaOH(aq) --> NaCl(aq) + H2O(l)

HCl and NaOH become ions in solution, so the ionic equation is:

- H+ + Cl− + Na+ + OH− --> Na+ + Cl− + H2O(l)

Since sodium and chloride ions are not involved in the reaction, the equation is:

- $H^+ + OH^- \longrightarrow H_2O(l)$: $\Delta H_r = -55.90$ kJ/mol

Krebs cycle

Also known as the citric acid cycle, the Krebs cycle is an important metabolic pathway in the series of pathways by which large molecules are broken down into energy for use by living organisms. It was named for its discoverer, Hans Adolf Krebs, who described the pathway in 1937. The cycle is the pathway in the metabolism of sugars. After a person eats a meal, for example, the food particles are converted to glucose and/or glycerol; they then undergo another pathway which further reduces the glucose/glycerol molecules to pyruvate, and then yet further to acetyl coenzyme A. That substance then enters the Krebs cycle where a complex series of reactions produce ATP (adenosine triphosphate), NAD (nicotinamide adenine dinucleotide), and NADP (nicotinamide adenine dinucleotide phosphate). A number of other pathways feed into the cycle at various points, providing necessary inputs and outputs for it and other pathways.

The citric acid (or Krebs) cycle takes place within the mitochondria of cells. It follows glycolysis and oxidative decarboxylation in the string of metabolic pathways through which sugars are converted to forms of energy usable by cells. It involves two rotations, or turns, through the cycle and begins when a molecule of acetyl coenzyme A (acetyl-CoA) forms citrate (a carboxylic acid also called citric acid) by reacting with a four-carbon carboxylate. A series of reactions converts the citrate back to a four-carbon carboxylate, releasing CO_2 in the process and producing 1 ATP, 3 NADH, and 3 FADH. Because the input of pyruvic acid involves two molecules, the cycle "turns" again doubling the products to 2 ATP, 6 NADH, and 6 FADH. Water and 4 CO_2 are expelled as by-products.

Henderson-Hasselbalch equation

The Henderson-Hasselbalch equation describes the impact of pH on conjugate acid-base mixtures in equilibrium. More specifically, it derives the pH of a buffer without first having to calculate equilibrium. It is meant to simplify the calculation of pH and buffer strength by using small, positive numbers for constants instead of very small exponents (such as 10^{-14}) by taking the logarithm of the relative values of dissociated ions and their parent molecules. These numbers are represented in the equation by pKa and pKb. In each use of pK_a below, assume that pKb could be substituted. The full equation is as follows:

$$pH = pK_a + log\frac{[A^-]}{[HA]}$$

where A- is the anion of any hydrogen based acid. If A- and HA are equal, the pH and pK_a will be equal because the log of 1 is 0.

Bases

Basic chemicals are usually in aqueous solution and have the following traits: a bitter taste; a soapy or slippery texture to the touch; the capacity to restore the blue color of litmus paper which had previously been turned red by an acid; the ability to produce salts in reaction with acids. "Alkali" is often used to describe bases. While acids yield hydrogen ions (H^+)when dissolved in solution, bases yield hydroxide ions (OH^-); the same models used to describe acids can be inverted and used to describe bases— Arrhenius, Brønsted-Lowry, and Lewis. Some nonmetal oxides (such as Na_2O) are classified as bases even though they do not contain hydroxides in their molecular form. However, these substances easily produce hydroxide ions when reacted with water, which is why they are classified as bases.

Neutralization reactions

Neutralization reactions are associated with acids and bases and are named as such due to the fact that the reaction cancels out, or neutralizes, the acidic and basic properties of the reactants. The foundation of neutralization reactions is the dissociation of water: H_2O Ý $H^+ + OH^-$; this ionization of water is the principal reaction of all aqueous acid-base reactions. According to the Arrhenius model, it is the only acid-base reaction (other theories are not so limiting). Neutralization reactions are exothermic and yield water and salts as in this example:

$$Na^+OH- + HCl^- \rightarrow Na^+Cl^- + H_2O.$$

The salts produced in neutralization reactions are simply the ionic compounds of the anion and cation that were already present in the acid and the base. Because the reactions occur in solution, the salts are not solids, though some will precipitate out of solution. The solid salts may be recovered by evaporating the water.

Buffering

Buffers are compounds, in solutions, which help to maintain the pH of that solution at an almost precise level despite other acids or bases that may be added. Solutions containing buffers are also sometimes called buffers, or are said to be "buffered." For example, if a solution containing a buffer has a pH of 5 (a fairly weak acid), the pH will not change appreciably if more acid—thus more dissociable H ions—are added to the mix. Buffers are usually made of a weak acid and the salt of that acid; the negative ions of the salts act as "ion traps" for the H+ which dissociate from the weak acid - they can't freely dissociate and change the pH of the solution. Buffer capacity is the upper limit of how much "protection" the buffer can provide; too much acid or base added to a solution will overwhelm the buffer.

Aprotic solvent, amphiprotic solvent, agua regia, and autoprotolysis

Aprotic solvents are those having neither acidic nor basic properties, e.g. having no dissociable H atoms. Pentane and toluene are aprotic solvents. By contrast, amphiprotic solvents exhibit both acidic and basic properties because they have dissociable H atoms. This term is synonymous with amphoteric and ampholytic. Water, ammonia, and ethanol are examples of amphiprotic solvents.

Aqua regia is a mixture of nitric acid and hydrochloric acid used in industrial applications such as the dissolution of gold. Autoprotolysis refers to the transfer, or dissociation and reassociation, of H atoms within the same substance. The autoionization of water is a typical example; H and O atoms in water molecules separate to form H^+ and OH^- ions, and then rejoin in the H_2O molecular form. This autoprotolysis exists in equilibrium: H_2O (l) Ý H^+(*aq*) + OH^-(*aq*)

[OH-], pH, and pOH

OH^- is the hydroxide ion associated with acids. Placing it in square brackets indicates a number of OH^- ions as a variable in equations. pH refers to the "power of Hydrogen" ions concentrated in solution; it is a measurement of the acidity or alkalinity of solution and ranges from 0 to 14. Neutral solutions have a pH of 7; solutions become more acidic as their pH decreases from 7 and more basic as their pH increases from 7. pH is derived as the logarithm of the concentration of H ions: pH = -log[$H^{+]}$ In this example, the value of [H+] = 1.0 x 10-7: pH = -log[H^+] = -log(1.0 x 10-7) = 5.27, an acidic solution. pOH, the "power of hydroxide ion" is another measurement of concentration in solutions. It is closely related to pH, as is the equation for calculating it: pOH = -log[$OH^{-]}$.

Common-ion effect

The common-ion effect describes the impact of adding substances to solution that have ions in common with the solute or solvent. Imagine a solution, H_2O + NaCl, in equilibrium. Adding more salt or more water to the solution causes the equilibrium to shift because the added material has ions in common with the solution. The common-ion effect allows predictions about pH because of the known and quantifiable impacts of changes in solution equilibria. Another solution in equilibrium is shown below:

$HC_2H_3O_2$(*aq*)ÝH+(aq) + $C_2H_3O_2$-(*aq*)

If we add iron acetate ($FeC_2H_3O_2$) to the solution, we can see that the iron and acetate will separate to form the ions Fe+ and $C_2H_3O_2$-. Because the number of acetate ions in the solution will have increased, the equilibrium will shift to the left and establish a new equilibrium as predicted by Le Châtelier's principle.

Titration

Titration is the process by which the amount of a given substance present in solution is determined. If a solution contains an unknown amount of acid, for instance, titration calculations can determine the unknown value by measuring the moles of base required to completely neutralize the acid. Titration can also work in reverse: the unknown quantity of a base in solution is found by adding measured amounts of acid until the base is neutralized. The substance whose amount is unknown is “being titrated;” the substance added in known amounts is the “titrant.” In addition to determining concentration, titration can also determine K_a and K_b by slightly modifying the calculations. Titration is typically shown on a titration curve; important changes in the angle of the curve correspond to changes in pH and when neutralization has been achieved, a point called the equivalence point, the pH of the solution will be 7 in the titration of strong acids with strong bases.

Acid-base indicator and acid dissociation constant

Acid-base indicators are substances—often in solution—which indicate changes in pH by changes in color. They are themselves weak acids.

The acid dissociation constant, K_a in equations, is also known as the acid ionization constant. It is a value expressing the point at which an acid dissociation reaction is at equilibrium. In other words, it measures the amount of acid that will break into H^+ and an anion before those ions begin to recombine into their parent molecular form. For instance, acetic acid dissociates as follows: $HC_2H_3O_2$(aq) Ý H^+(aq) + $C_2H_3O_2^-$(aq). The equilibrium constant of this equilibrium reaction may be determined by the following equation where square brackets indicate numbers of particles:

$$K_a = \frac{[H^+][C_2H_3O_2^-](aq)}{[HC_2H_3O_2]}$$

This equation is the specific application of a general formula for calculating the value of K_a (below). The formula dictates that the greater the value of HA, the greater the dissociation (or strength) of the acid and the value of K_a.

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Base constant and water ionization constant

In addition to the acid dissociation constant (K_a) discussed above, the base and water ionization constants are some of the most important constants in acid-base and solution chemistry. The capital letter K is the chemistry symbol for equilibrium constant and the superscript indicates which constant is being measured. K_b, the base ionization (or dissociation) constant, measures a base’s strength in reaction with hydronium ions to form the conjugate acid of the base. The value of K_b is inversely related to basicity: the higher the value, the weaker the base.

The water ionization constant, K_w, is the foundation of all equilibrium constants because other constants are measured in comparison to water. Because so few of the H atoms in pure H_2O dissociate, K_w is so small (10-14) as to be virtually zero. This provides a known and reliable baseline against which to measure other constants and equilibria.

Oxidation and Reduction

Redox chemistry

Redox chemistry is the branch of chemistry dedicated to the study of redox, or oxidation-reduction reactions. These reactions are also called displacement reactions because they are characterized by the replacement of one or more atoms in a compound with atom(s) of a different element. Redox reactions involve two halves—an oxidation occurs when one of the reactants loses electrons thus becoming positively charged. The partner in the reaction becomes reduced when it gains electrons and becomes negatively charged. Due to the transfer of electrons, all redox reactions involve changing the oxidation numbers of the reactants. The change in number is a more precise measurement of redox reactions, as the actual transfer of electrons may not occur; the charges are still generated because of the greater electronegativity of one of the reactants which draws its partner reactant's electrons closer to its nucleus.

Nonmetals

When nonmetal elements act as oxidizing agents, the nonmetal is reduced to a monatomic anion or a protonated form. The monatomic nonmetal anions and their protonated forms normally show the lowest oxidation number possible which allows them to transfer electrons back to other materials, or act as reducing agents. Nonmetals can be divided into two groups which are based on their redox properties. First, there are the very electronegative nonmetals. These nonmetals have an electronegativity of > 2.8 and are good oxidizing agents. Although O2 is a strong oxidizing agent, it has oxidizing reactions that are normally slow kinetically. Many flammable materials exist in the presence of air until a reaction is started by a spark or flame. The second category of nonmetals is simply, electronegative nonmetals. These have electronegatives between 1.9 and 2.8. These are elements that have few uses in the laboratory as oxidizing agents.

Chemical changes

Experience and observation are necessary in order to identify chemical changes. Signs to look for include a change in color, the formation of solids, the formation of bubbles, heat and/or flame is produced or heat is absorbed. Some changes that are not chemical include the freezing or boiling of water, which are physical changes. Adding a solid to the water such as sodium chloride and it dissolving is a chemical change. This is even though the sodium chloride will disappear because if most of the water evaporates then most of the sodium chloride is recovered. The important factor in determining a chemical change is the formation of a new substance.

Oxidation of iron

Iron has the ability to rust under the right conditions which is a chemical property. Rusting is a slow chemical change. This is because rust is an iron oxide. It has different properties than iron metal. This can be illustrated by imagining steel wool burning in a fast reaction with oxygen, then contrast it with the slow rusting of iron which is also oxygenated. In iron, the element, only the atoms of iron are contacting one another. In the element oxygen, each oxygen atom is joined together to form a diatomic molecule. There is a rearranging of the atoms and molecules so that two iron atoms are combined with three oxygen atoms. This forms a completely new compound.

Proteins

Proteins can also be metabolized for use as a source of energy, but this is merely one catabolic pathway associated with them. Because proteins are constructed of long chains of amino acids, reactions early in the process are of particular interest. One is the oxidation of cysteine, a sulfur-containing amino acid. Two cysteine molecules combine with oxygen and fuse to form an S—S disulfide bond, creating one molecule of cystine. This reaction is important in determining the unique structure of some proteins. Peptides are groups of two or more linked amino acids. Peptide formation occurs during protein synthesis when a carboxyl group reacts with an amine of a second amino acid to release water and form an amide bond which links the two amino acids into a peptide. The amide bond is called a peptide bond; many such reactions can produce long chains (polypeptides) that eventually become proteins.

Protein hydrolysis and denaturation

Protein hydrolysis is one of the processes within the metabolic pathway that produces energy for use by living cells. Proteins may be hydrolyzed by acids or bases or by water due to its highly polar molecules. Long protein chains are broken into smaller peptides and, with increased time, temperature, or acidity/basicity, the peptides can be reduced to individual amino acids. Hydrolysis reactions in digestion are catalyzed by enzymes present in animals' digestive systems. Denaturation is the process by which the orderly structure of natural proteins is broken through exposure to conditions such as high temperatures and certain compounds. The protein's folds and helical structures straighten and twist into random forms, making the protein inactive. Denaturation of toxic proteins explains why boiling water to kill bacteria makes it safe to drink—the bacteria no longer produce harmful proteins and the present proteins become denatured.

Protein synthesis

One of the most fundamental processes for life, protein synthesis, is a complex series of reactions involving large compounds. The synthesis occurs at the surface of ribosomal RNA (rRNA) particles suspended in the cytoplasm of cells. Amino acids bonded to transfer RNA (tRNA) are brought to the sites and the translation of codons into a sequence of amino acids begins if the required compounds are present. If all requirements are met, a series of enzyme-catalyzed steps begins to form long peptide chains. The process begins with the initiation step and concludes with termination, at which point the protein is separated from the rRNA. Most proteins are not yet in the form needed by the cell and are modified by more complex reactions to produce the needed protein.

Redox reactions

Electron transfer, or redox reactions, occurs when electrons move from one atom to another, changing the charge of the ion. Because the charge has changed, the oxidation number also changes. Oxidation is an important class of redox reactions in which the oxidation number increases. Commonly, any reaction in which oxygen combines with other substances is oxidation. Rusting iron and burning wood are both examples of oxidation. Precipitation, or ion combination reactions, occurs when positive and negative compounds in solution combine to form an insoluble ionic compound. Acid-base reactions occur when an acid reacts with a base and an ion of hydrogen transfers to the base. Polymerization reactions occur when simple molecules, also called monomers, combine to form complex molecules, or polymers.

Oxidizing agent, reduction agent, and rate-determining step

An oxidizing agent is the reactant in oxidation reactions which gains electrons, causing oxidation of the other reactant(s). Peroxides, iodine and other halogens, and sulfoxides are common oxidizing agents. A reduction, or reducing agent, is the reactant oxidized in an oxidation reaction; it loses electrons to the oxidizing agent. In rusting iron, a common oxidation reaction, iron is the reducing agent, losing electrons to oxygen. In multi-step reactions, each of the steps has different reaction rates which are often very different. The rate-determining step is the slowest portion of such reactions. Because the reaction can only go as fast as its slowest step, that step determines overall reaction rate.

Oxidation number

The oxidation number, or oxidation state, is a representation of the combination state of an element. Because elements produce different ions depending upon the compounds in which they occur and the types of bonds in which they are involved, the oxidation number identifies the ion in question. If the atom is isolated or in its uncombined state, the oxidation number is 0. The oxidation state of any given atom is given by the charge of the ion. For example, when fluorine occurs in ionic compounds, it is the fluoride ion, F-. When in covalent bonds, fluorine shares a single pair of electrons; the oxidation number -1 reflects this behavior. A single fluorine atom, uncombined with other elements has an oxidation number of 0. A calcium ion, Ca^{2+}, has an oxidation number of +2. The oxidation number enables accurate predictions for chemical reactions.

Organic Chemistry

Metabolism and homeostasis

Metabolism and homeostasis, present in all living things, are two essential functions for the maintenance of life.

Metabolism refers to the cell's ability to extract energy from its environment and use that energy for cell function, growth, and maintenance. The cell also uses this energy to reproduce itself as directed by DNA. Metabolism is a conversion of energy from external sources to fuel cell life. Photosynthesis is a prime example of cell metabolism. Light energy is converted to a specialized molecule, adenosine triphosphate (ATP), that the cell uses as a basic building block.

Homeostasis is the ability of a living cell to maintain an internal environment favorable to life. Internal systems ensure that a life form adjusts to changing conditions and survives. Examples are self-regulating thermostats in the body, which keep temperatures in a healthy range. Both simple and complex organisms have these internal controls that maintain homeostasis. Without this capacity, living beings could not survive changing conditions.

Organic synthesis and retrosynthesis

Organic synthesis is the chemical process by which organic and biomolecules are constructed. It is a large and growing branch of organic chemistry. Synthesis may be linear, in which complex molecules are the product of several intermediate steps in multi-step reactions; they may also be convergent, in which necessary parts of the molecule are created in separate reactions then brought together to from the desired product. Retrosynthesis refers to a methodology for the study of organic synthesis in which analysis begins with the products and works backwards, using known laws and steps to arrive at reactants. The Wöhler synthesis is an early example of attempts to understand organic synthesis. Organic synthesis occurs naturally in biological processes and artificially in laboratories and industry, such as in the production of plastics and synthetic fibers.

Functional groups

"Functional groups" relate to organic chemistry; they are atoms or groups of atoms (two to four) within molecules which give the molecules their distinctive chemical reactions. Regardless of the other atoms surrounding the functional groups in a molecule, they will always react in the same or similar ways. Functional groups are the basis of nomenclature for organic compounds and are an important class of organic chemicals. Because their reactions are consistent and predictable, they are often the "pivot points" for discussions in organic chemistry. The presence of a functional group in a molecule defines that molecule as belonging to a specific chemical class. For example, organic molecules containing a hydroxyl group are classified as Alcohols. Some common functional groups are: Hydroxyl, Methyl, Alkyne, Amide, primary Amine, secondary Amine, tertiary Amine, Azo, Nitrite, Nitro, Nitroso, Pyridyl, Carboxyl, Aldehyde, and Ketone.

Chemical makeup

The following is a list of the chemical makeup of common functional groups:

- Hydroxyl: Characteristic of alcohols, the hydroxyl group is composed of an ion of OH single bonded to a hydrocarbon.
- Carbonyl: a C double bonded to an O forms a carbonyl, visually represented: C=O, where the double line illustrates the double bond. Carbonyls with other atoms or simple molecules make up an important functional group including aldehydes, ketones, and esters.
- Carboxyl: A carbonyl that has reacted with an oxidizing agent and acquires an extra O is a carboxyl. In compounds they form an important class of organic acids called carboxylic acids.
- Amine: A base ammonia molecule (NH_3) in which one of the hydrogens has been replaced by an alkyl, forming $R—NH_2$, where R represents the alkyl.
- Esters: The product of a carboxyl's reaction with alcohol. The OH of alcohol and the H of the carboxyl condense to form water, leaving an ester: R—COO—R, where R represents any hydrocarbon to which the ester is bonded.

Classification of organic reactions

There are enormous numbers of known compounds and reactions important to organic chemistry. Classifying them is a necessary step in understanding and organizing them into usable information. The following categories are useful in such classification, although it is important to remember that each category contains various subcategories that further define the specific reactions:

- Structural change. These reactions include those in which reactants undergo structural changes. Four subclasses are in this category: addition, elimination, substitution, and rearrangement.
- Reaction type, including acid-base, oxidation, and reduction reactions.
- Functional group. Many functional groups exist in organic chemistry, each with characteristic reactions. Just as organic nomenclature is based on which functional groups are present, reactions are also often determined by the functional group. Dehydration reactions, for instance, are typical of alcohols; while combustion reactions are typical of alkanes.

Wöhler synthesis

The Wöhler synthesis is an important reaction in the study of organic chemistry. Friedrich Wöhler synthesized urea, an organic compound, from inorganic elements (ammonium cyanate) in 1828, providing a key to the modern study of organic chemistry. Prior to Wöhler, chemists believed that organic compounds could not be synthesized and were animated by something called the "living force."

Alcohols

Composition and nomenclature

Alcohols are characterized by the functional group hydroxyl—OH—bonded to the parent hydrocarbon molecule. Alcohol is highly reactive because of the covalent bonds at both the C—O and O—H sites, which leave more "unattached" electrons free to engage in chemical reactions. This high reactivity accounts for the many uses of many types of alcohols, and its dangerous attributes. Alcohols are named by first identifying the parent hydrocarbon, usually an alkyl, and indicating the hydroxyl function by adding the –ol suffix, as in ethanol and propanol. Organic compounds may contain more than one hydroxyl group, these are called polyhydroxy alcohols. The reactions most often associated with alcohols have to do with the bonds of the –OH group. In some classes of alcohol reactions, the bond between the O and H is broken, while in others the bond between the O and C is broken, producing alkyl halides.

Measurement and Data Processing

Measuring, organizing, and classifying data

Measuring data is a crucial part of the scientific process. Measurements are most useful if they are quantified—expressed in numbers. Measuring is the process of determining variables such as time, space, and temperature of objects and processes in precise numbers. The metric system is the universal standard of measurement in science.

Data must be organized in a practical, useful manner to be valuable. Scientists use graphs, charts, tables, and other organizational tools to make data more useful.

Data must then be classified—grouped into organizational schemes for easy access and use. These schemes attempt to organize the maximum amount of useful data in a format that scientists can use. Although these steps may be less glamorous than other areas of science, they are essential.

Verification and confirmation of data

A critical distinction should be made between confirmation and verification of scientific data. *Verification* establishes once and for all the truth of the statement. *Confirmation* is the testing of claims to see how true they are. For a claim to be testable, an experiment must be devised to ensure the validity of the results. A claim can only be confirmed when we know the conditions for verification. A claim confirmation is always relative to the testing procedures. Test results must always be objective and observable. Actually, no factual claim can ever be verified because there is always the possibility of new evidence appearing that proves the claim false. A scientific law must also be confirmed by making predictions based on unbiased, observable data.

Practice Test

Practice Questions

1. Which substance is most likely to be a solid at STP?
   - a. Kr
   - b. Na
   - c. NH_3
   - d. Xe

2.Which of the following tend to increase the melting point of a solid?
   - *I. Increasing molecular weight*
   - *II. Decreasing polarity*
   - *III. Increasing* surface area
   - a. I and II
   - b. II
   - c. III
   - d. I and III

3. A gas at constant volume is cooled. Which statement about the gas must be true?
   - a. The kinetic energy of the gas molecules has decreased.
   - b. The gas has condensed to a liquid.
   - c. The weight of the gas has decreased.
   - d. The density of the gas has increased.

4. A weather balloon is filled with 1000 mol of He gas at 25 °C and 101 kPa of pressure. What is the volume of the weather balloon?
   - a. 24518 m3
   - b. 24.5 m3
   - c. 2 m3
   - d. 245 m^3

5. One mole of oxygen gas and two moles of hydrogen are combined in a sealed container at STP. Which of the following statements is true?
   - a. The mass of hydrogen gas is greater than the mass of oxygen.
   - b. The volume of hydrogen is greater than the volume of oxygen.
   - c. The hydrogen and oxygen will react to produce 2 mol of water.
   - d. The partial pressure of hydrogen is greater than the partial pressure of oxygen.

6. Graham's law is best used to determine what relationship between two different materials?
   - a. pressure and volume
   - b. volume and temperature
   - c. mass and diffusion rate
   - d. Diffusion rate and temperature

7. Which is the correct order of increasing intermolecular attractive forces?
   a. Dipole-dipole<ionic<hydrogen bonding<London dispersion
   b. Ionic<dipole-dipole<London dispersion<hydrogen bonding
   c. Hydrogen bonding<London dispersion<ionic<dipole-dipole
   d. London dispersion<dipole-dipole<hydrogen bonding<ionic

8. One mole of an ideal gas is compressed to 10 L and heated to 25 °C. What is the pressure of the gas?
   a. 2.4 KPa
   b. 2.4 atm
   c. 0.2 atm
   d. 0.2 KPa

9. A 10 L cylinder contains 4 moles of oxygen, 3 moles of nitrogen and 7 moles of neon. The temperature of the cylinder is increased from 20 °C to 40 °C. Determine the partial pressure of neon in the cylinder as a percentage of the whole.
   a. 50%
   b. 70%
   c. 90%
   d. 40%

10. Three liquids, X, Y and Z are placed in separate flasks, each of which is suspended in a water bath at 75 °C. The boiling points of each liquid are
      X, 273 K
      Y, 340 K
      Z, 360 K
Which of the three liquids will begin to boil after warming to 75 °C?
   a. X, Y, and Z
   b. X and Z
   c. X and Y
   d. Y and Z

11. Which of the following statements is true about the physical properties of liquids and gases?
      I. Liquids and gases are both compressible
      II. Liquids flow, but gases do not
      III. Liquids flow, and gases are incompressible
      IV. Liquids flow and gases are compressible
      V. Gases flow and liquids are incompressible
   a. I and III
   b. II and IV
   c. III and V
   d. IV and V

12. Which of the following statements **generally** describes the trend of electronegativity considering the Periodic Table of the Elements?
   a. Electronegativity increases going from left to right and from top to bottom
   b. Electronegativity increases going from right to left and from bottom to top
   c. Electronegativity increases going from left to right and from bottom to top
   d. Electronegativity increases going from right to left and from top to bottom

13. Gas X is in a cylinder at 1 atm of pressure and has a volume of 10 L at 0° C. Gas X spontaneously decomposes to gas Y, according to the equation

$$X \longrightarrow 3Y$$

The temperature in the cylinder remains the same during the reaction. What is the pressure in the cylinder now?

a. 1 atm
b. 3 atm
c. 4 atm
d. Cannot be determined

14. 1 mole of water and 1 mole of argon are in a cylinder at 110 °C and 1 atm of pressure. The temperature of the cylinder is reduced to -5 °C. Which statement about the contents of the cylinder is most accurate?

a. The pressure in the cylinder is decreased, and the partial pressure of argon is less than that of water.
b. The pressure in the cylinder is about the same, and the partial pressure of water is less than that of argon.
c. The pressure in the cylinder is decreased, and the partial pressure of water is much less than that of argon.
d. The pressure in the cylinder is decreased and the partial pressure of water is the same as argon.

15. A solid is heated until it melts. Which of the following is true about the solid melting?

a. ΔH is positive, and ΔS is positive
b. ΔH is negative and ΔS is positive
c. ΔH is positive and ΔS is negative
d. ΔH is negative and ΔS is negative

16. A liquid is held at its freezing point and slowly allowed to solidify. Which of the following statements about this event are true?

a. During freezing, the temperature of the material decreases
b. While freezing, heat is given off by the material
c. During freezing, heat is absorbed by the material
d. During freezing, the temperature of the material increases

17. A liquid is heated from 50 °C to 80 °C. Which of the following statements is generally true about the solubility of solids and gases in the liquid?

a. The solubility of solids will increase and the solubility of gases will decrease
b. The solubility of solids will decrease and the solubility of gases will increase
c. The solubility of solids will increase and the solubility of gases will increase
d. The solubility of solids will decrease and the solubility of gases will decrease

18. 100 g of H_3PO_4 is dissolved in water, producing 400 mL of solution. What is the normality of the solution?

a. 2.55 N
b. 1.02 N
c. 7.65 N
d. 0.25 N

19. Silver nitrate ($AgNO_3$) is dissolved in water. One drop of an aqueous solution containing NaCl is added and almost instantly, a white milky precipitate forms. What is the precipitate?
   a. NaCl
   b. $NaNO_3$
   c. $AgNO_3$
   d. AgCl

20. 100 g of ethanol C_2H_6O is dissolved in 100 g of water. The final solution has a volume of 0.2 L. What is the density of the resulting solution?
   a. 0.5 g/mL
   b. 1 g/mL
   c. 46 g/mL
   d. 40 g/mL

21. 100 mL of a 0.1 M solution of NaOH is neutralized to pH 7 with H_2SO_4. How many grams of H_2SO_4 are required to achieve this neutralization?
   a. 4.9 g
   b. 0.98 g
   c. 9.8 g
   d. 0.49 g

22. Comparing pure water and a 1 M aqueous solution of NaCl, both at 1 atm of pressure, which of the following statements is most accurate?
   a. The pure water will boil at a higher temperature, and be less conductive
   b. The pure water will boil at a lower temperature and be less conductive
   c. The pure water will boil at a lower temperature and be more conductive
   d. The pure water boil at the same temperature and be more conductive

23. Place the following in the correct order of increasing acidity.
   a. HCl<HF<HI<HBr
   b. HCl<HBr<HI<HF
   c. HI<HBr<HCl<HF
   d. HF<HCl<HBr<HI

24. Place the following in the correct order of increasing solubility in water.
   a. Butanol<ethanol<octane<NaCl
   b. Ethanol<NaCl<octane<butanol
   c. NaCl<octane<butanol<ethanol
   d. Octane<butanol<ethanol<NaCl

25. 50 grams of acetic acid $C_2H_4O_2$ are dissolved in 200 g of water. Calculate the weight % and mole fraction of the acetic acid in the solution.
   a. 20%, 0.069
   b. 0.069%, 0.20
   c. 25%, 0.075
   d. 20%, 0.075

26. Ammonium Phosphate $(NH_4)_3PO_4$ is a strong electrolyte. What will be the concentration of all the ions in a 0.9 M solution of ammonium phosphate?
a. 0.9 M NH_4+, 0.9 M PO_4^{3-}
b. 0.3 M NH_4+, 0.9 M PO_4^{3-}
c. 2.7 M NH_4+, 0.9 M PO_4^{3-}
d. 2.7 M NH_4^+, 2.7 M PO_4^{3-}

27. Which of the following represents the correct increasing order of acidity?
a. $CH_3COOH<CH_3OH<CH_3CH_3<HCl$
b. $CH_3CH_3<CH_3OH<CH_3COOH<HCl$
c. $CH_3CH_3<CH_3COOH<CH_3OH<HCl$
d. $CH_3OH<CH_3CH_3<HCl<CH_3COOH$

28. One liter of a 0.02 M solution of methanol in water is prepared. What is the mass of methanol in the solution, and what is the approximate molality of methanol?
a. 0.64 g, 0.02 m
b. 0.32 g, 0.01 m
c. 0.64 g, 0.03 m
d. 0.32 g, 0.02 m

29. A 1 M solution of NaCl (A) and a 0.5 M solution of NaCl (B) are joined together by a semi permeable membrane. What, if anything, is likely to happen between the two solutions?
a. No change, the solvents and solutes are the same in each
b. Water will migrate from A to B
c. NaCl will migrate from A to B and water will migrate from B to A.
d. Water will migrate from B to A.

30. Which of the following radioactive emissions results in an increase in atomic number?
a. Alpha
b. Beta
c. Gamma
d. Neutron

31. A material has a half life of 2 years. If you started with 1 kg of the material, how much would be left after 8 years?
a. 1 kg
b. 0.5 kg
c. 0.06 kg
d. 0.12 kg

32. C-14 has a half life of 5730 years. If you started with 1 mg of C-14 today, how much would be left in 20,000 years?
a. 0.06 mg
b. 0.07 mg
c. 0.11 mg
d. 0.09 mg

33. The best way to separate isotopes of the same element is to exploit:
  - a. Differences in chemical reactivity
  - b. Differences in reduction potential
  - c. Differences in toxicity
  - d. Differences in mass

34. Nuclear chain reactions, such as the one that is exploited in nuclear power plants, are propagated by what subatomic particle(s)?
  - a. Protons
  - b. Neutrons
  - c. Electrons
  - d. Neutrons and protons

35. Which of the following statements about radioactive decay is true?
  - a. The sum of the mass of the daughter particles is less than that of the parent nucleus
  - b. The sum of the mass of the daughter particles is greater than that of the parent nucleus
  - c. The sum of the mass of the daughter particles is equal to that of the parent nucleus
  - d. The sum of the mass of the daughter particles cannot be accurately measured

36. Determine the number of neutrons, protons and electrons in ^{238}U.
  - a. 238, 92, 238
  - b. 92, 146, 146
  - c. 146, 92, 92
  - d. 92, 92, 146

37. An alpha particle consists of
  - a. Two electrons and two protons
  - b. Two electrons and two neutrons
  - c. Four neutrons
  - d. Two protons and two neutrons

38. Describe the correct outer shell electronic arrangement of phosphorous.
  - a. $4s^2\ 4p^3$
  - b. $3s^2\ 3p^3$
  - c. $2s^2\ 3p^3$
  - d. $2s^2\ 2p^3$

39. Hund's rule regarding electronic configuration states:
  - a. Electrons in the same orbital must have an opposite spin
  - b. Electrons must fill lower energy orbitals before filling higher energy orbitals
  - c. Electrons must populate empty orbitals of equal energy before filling occupied orbitals
  - d. Electrons must have the same nuclear spin as the nucleus

40. Arrange the following elements in order of increasing atomic radius:
  - a. K<Zn<Fe<As<Kr
  - b. K<Fe<Zn<Kr<As
  - c. Kr<As<Fe<K<Zn
  - d. Kr<As<Zn<Fe<K

41.When a solid is heated and transforms directly to the gaseous phases, this process is called:

a. sublimation
b. fusion
c. diffusion
d. condensation

42. Determine the oxidation states of each of the elements in $KMnO_4$:

a. K^{+1}, Mn^{+7}, O^{-8}
b. K^{-1},Mn^{+7}, O^{-2}
c. K^{+1}, Mn^{+3}, O^{-4}
d. K^{+1}, Mn^{+7},O^{-2}

43. Place the following elements in order of decreasing electronegativity:

N, As, Bi, P, Sb

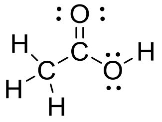
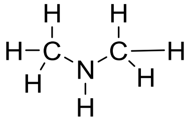
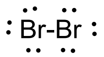
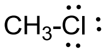
a. As>Bi>N>P>Sb
b. N>P>As>Sb>Bi
c. Bi>Sb>As>P>N
d. P>N>As>Sb>Bi

44. Arrange the following compounds from most polar to least polar:

F_2, CH_3CH_2Cl, NaCl, CH_3OH

a. NaCl>CH_3OH> CH_3CH_2Cl> F_2
b. F_2> NaCl> CH_3OH>CH_3CH_2Cl
c. CH_3OH>NaCl>F_2>CH_3CH_2Cl
d. NaCl>F_2>CH_3OH>CH_3CH_2Cl

45. Which of the following is an incorrect Lewis structure?

I.

II.

III.

IV.


a. I
b. II
c. III
d. IV

46. Which bond has the shortest length?

a. sp^2
b. sp^3
c. sp
d. pi

47. Resonance structures can be defined as:

a. Two or more structures that have different atoms bound to different atoms
b. Two structures that have a similar structure but different formula
c. Two or more structures that have the same formula, but are different in shape
d. Two or more structures that differ only in the arrangement of electrons in the structures

48. Atoms that are sp^2 hybridized will have what sort of hybrid orbital geometry around them?

a. Tetrahedral
b. Trigonal planar
c. Linear
d. Angled

49. What is the chemical composition of ammonium sulfate?

a. N 21%, H 3%, S 24%, O 32%
b. N 10%, H 6%, S 24%, O 60%
c. N 10%, H 4%, S 12%, O 74%
d. N 21%, H 6%, S 24%, O 48%

50. What is the correct IUPAC name of the compound Fe_2O_3?

a. Iron (I) oxide
b. Iron (II) oxide
c. Iron (III) oxide
d. Iron (IV) oxide

51. Balance the following reaction between sulfuric acid and aluminum hydroxide by filling in the correct stoichiometric values for each chemical.

$\_\ H_2SO_4 + \_\ Al(OH)_3 \rightarrow \_\ Al_2(SO_4)_3 + \_\ H_2O$

a. 3, 2, 1, 6
b. 2, 3, 1, 3
c. 3, 3, 2, 6
d. 1, 2, 1, 4

52. Calculate the mass of water produced from the reaction of 1 kg of n-heptane with oxygen.

n-heptane (1 kg) $+ 11\ O_2 \rightarrow 7\ CO_2 + 8\ H_2O$

a. 144 g
b. 8 kg
c. 800 g
d. 1.4 kg

53. Magnesium metal is reacted with hydrobromic acid according to the following equation:

$Mg + 2HBr \rightarrow MgBr_2 + H_2$

If 100 g of Mg is reacted with 100 g of HBr, which statement about the reaction is true?

a. Mg is the limiting reagent
b. HBr is the excess reagent
c. Mg is the excess reagent
d. 100 g of $MgBr_2$ will be produced

54. Methane gas is burned in pure oxygen at 200 °C and 1 atm of pressure to produce CO_2 and H_2O according to the equation

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

If 10 L of methane gas were burned, and the final temperature and pressure remained the same, how many liters of gaseous products are produced by the reaction?

a. 10 L
b. 20 L
c. 30 L
d. 40 L

55. The overall reaction A→D can be described by the following equation:

$$A \xrightarrow{\text{fast}} B \xrightarrow{\text{slow}} C \xrightarrow{\text{fast}} D$$

What would be the rate law for the overall reaction of A to D?

a. Rate = $k[D]/[A]$
b. Rate = $k[B]$
c. Rate = $[B]$
d. Rate = $k[C]/[B]$

56. How many electrons are in the atom $^{45}_{20}\text{Ca}$?

a. 20
b. 45
c. 65
d. 25

57. For the reaction $CO_2(g) + H_2(g) \rightarrow CO(g) + H_2O(l)$, which of the following will occur if the pressure of the reaction is increased?

a. The reaction rate will increase
b. The reaction rate will decrease
c. The reaction equilibrium will shift to the right
d. The reaction equilibrium will shift to the left

58. For the gas phase reaction $CH_4 + 4Cl_2 \rightarrow CCl_4 + 4HCl$, what would be the equilibrium expression K_{eq} for this reaction?

a. $[CH_4][Cl_2] / [CCl_4][4HCl]$
b. $[CH_4][Cl_2] / [CCl_4][HCl]^4$
c. $[4Cl][CCl_4\}/[CH_4][4HCl]$
d. $[CCl_4][HCl]^4/ [CH_4][Cl_2]^4$

59. Adding a catalyst to a reaction will do which of the following to that reaction:

a. Shift the reaction equilibrium towards the products
b. Increase the temperature of the reaction
c. Decrease the energy of activation for the reaction
d. Increase the purity of the reaction products

60. 10 g of salt XY (MW = 100 g/mol) is added to 1 liter of water with stirring. The salt dissociates into ions X^+ and Y^-. After equilibrium is established, the undissolved portion of the salt was removed by filtration, weighed, and found to be 9.5 g. What is the K_{sp} for this salt?

a. 5×10^{-2}
b. 5 x 10-3
c. 1 x10-2
d. 2.5×10^{-5}

61. Which of the following are considered Lewis acids?

I. H_2SO_4
II. $AlCl_3$
III. PCl_3
IV. $FeCl_3$

a. II and IV
b. II and III
c. I and IV
d. I and II

62. Place the following in the correct order of increasing acidity:

H_3PO_4, HF, HCl, H_2O, NH_3

a. $H_3PO4<H_2O<NH_3<HF<HCl$
b. $NH_3<H_2O<HF<H_3PO_4<HCl$
c. $H_2O<NH_3<HF<H_3PO_4<HCl$
d. $NH_3<H_2O<HF<HCl<H_3PO_4$

63. The pka for ethanol (CH_3CH_2OH) is approximately 16. The pka for acetic acid (CH_3COOH) is about 4. The difference can be explained by:

a. Resonance stabilization
b. Electronegativity differences
c. Molecular weight differences
d. Molecular size differences

64. What will be the pH of 2 L of a 0.1 M aqueous solution of HCl?

a. 2
b. -1
c. 1
d. 0.05

65. What is the pH of a buffer containing 0.2 M NaOAc and 0.1 M HOAc? The pka of acetic acid is 4.75.

a. 4
b. 5
c. 6
d. 7

66. 50 mL of 1 M H_2SO_4 is added to an aqueous solution containing 4 g of NaOH. What will the final pH of the resulting solution be?

- a. 5
- b. 6
- c. 7
- d. 9

67. To make a good buffering system in the pH range of 5-9, which acid/base combinations would likely work the best?

- a. HCl/NaOH
- b. $HNO_3/NaNO_3$
- c. $H_2SO_4/NaHSO_4$
- d. NaH_2PO_4/Na_2HPO_4

68. For the conversion of water into steam, which of the following is true?

- a. $\Delta T=0, \Delta S>0$
- b. $\Delta T>0, \Delta S = 0$
- c. $\Delta T =0, \Delta S <0$
- d. $\Delta T >0, \Delta S >0$

69. 100 g of NH_3 are cooled from 100 °C to 25 °C. What is the heat change for this transition? The heat capacity of ammonia gas is 35.1 J/(mol) (°K)

- a. -263KJ
- b. 15.5 KJ
- c. -15.5KJ
- d. 263 KJ

70. Determine the heat of combustion for the following reaction:

Propane + 5 $O_2 \rightarrow 3 CO_2 + 4 H_2O$

The standard heats of formation for propane, CO_2 and water are -103.8 KJ/mol, -393.5 KJ/mol and -285.8 KJ/mol respectively.

- a. -2220 KJ/mol
- b. -2323.7 kJ/mol
- c. 2220 KJ/mol
- d. 2323.7 KJ/mol

71. Which of the following reactions produces products with higher entropy than the starting materials?

I. Glucose (s) + water $\rightarrow$glucose (aq)
II. $4Al (s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$
III. Br_2 + light$\rightarrow$2 Br
IV. Ice $\rightarrow$water vapor

- a. II, III
- b. I, II
- c. I, III
- d. I, III, IV

72. A 1kg block each of iron, lead and nickel are heated from 20 °C to 30 °C. Which of the following statements about the blocks is true?
   a. The lead will heat faster than the iron and the nickel.
   b. The iron required more heat to reach 30 °C than the nickel or lead.
   c. All three blocks required a different amount of heat to reach 30 °C.
   d. The iron required more time to reach 30 °C.

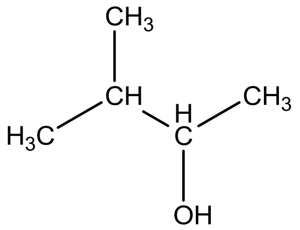
73. In the reaction $Pb + H_2SO_4 + H_2O \rightarrow PbSO_4 + H_2 + H_2O$
   a. Lead is reduced and hydrogen is oxidized
   b. Lead is oxidized and hydrogen is oxidized
   c. Lead is reduced and sulfate is oxidized
   d. Lead is oxidized and hydrogen is reduced

74. Which of the following elements would likely be good reducing agents?
   a. Br_2
   b. N_2
   c. Na
   d. Ne

75. Molten magnesium chloride is electrolyzed. The products formed from this reaction are:
   a. Mg(0) at the anode and Cl- at the cathode
   b. Mg2+ at the anode and Cl- at the cathode
   c. Mg (0) at the cathode and Cl_2 at the anode
   d. Mg(0) at the anode and Cl_2 at the cathode

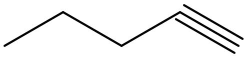
76. The transformation of diamond to graphite has a $-\Delta G$. Which of the following is true?
   a. The reaction is spontaneous and occurs rapidly at room temperature
   b. The reaction is not spontaneous and occurs slowly at room temperature
   c. The reaction is not spontaneous and does not occur at room temperature
   d. The reaction is spontaneous and occurs slowly at room temperature

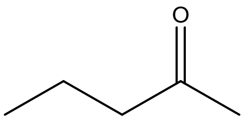
77. What would be the correct IUPAC name for the following compound?

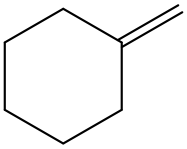


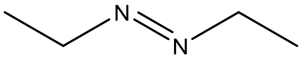
   a. 3-methyl-2-butanol
   b. 2-methyl-3-butanol
   c. 3,3-dimethyl-2-propanol
   d. 2-Hydroxy-3-methyl butane

78. Which of the following molecules are alkenes?

I. 

II. 

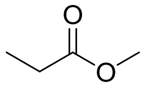
III. 

IV. 

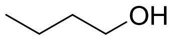
a. I
b. II
c. III
d. IV

79. What is the oxidation state of the carbon atom in a carboxylic acid functional group?
a. 4+
b. 3+
c. 2-
d. 3^-

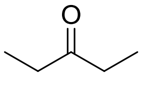
80. Which of the following molecules is named correctly?

A. 

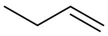
methyl propionoate

B. 

1-propanol

C. 

3-propanoic acid

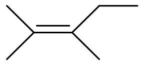
D. 

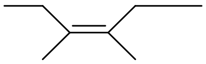
3-butene

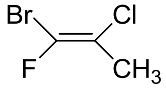
81. Which scientist was responsible for developing the modern periodic table?
a. Faraday
b. Einstein
c. Hess
d. Mendeleev

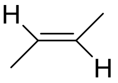
82. Two different molecules can be isomers of each other if:
a. They have the same functional groups
b. They have the same oxidation state
c. They have the same molecular weight
d. They have the same chemical formula

83. Which of the following molecules are cis alkenes?

I.


II.


III.


IV.


- a. I, II
- b. II, III
- c. III, IV
- d. I, IV

84. What would be the best analytical tool for determining the chemical structure of an organic compound?

- a. NMR
- b. HPLC
- c. IR
- d. Mass spec

85. Proteins are made up of which of the following repeating subunits?

- a. Sugars
- b. Triglycerides
- c. Amino acids
- d. Nucleic acids

86. The precision of a number of data points refers to:

- a. How accurate the data is
- b. How many errors the data contains
- c. How close the data points are to the mean of the data
- d. How close the actual data is to the predicted result

87. The density of a material refers to:

- a. Mass per volume
- b. Mass per mole
- c. Molecular weight per volume
- d. Moles per volume

88. Which of the following types of chemicals are considered generally unsafe to store together?

I. Liquids and solids
II. Acids and bases
III. Reducing agents and oxidizing agents
IV. Metals and salts

- a. I, II
- b. II, III
- c. III, IV
- d. I, IV

89. Which statement about the impact of chemistry on society is not true?

a. Fluoridation of water has had no effect on the rate of cavities as compared to unfluoridated water
b. Chemical fertilizers have tremendously increased food production per acre in the U.S.
c. Chemistry played a central role in the development of nuclear weapons
d. Use of catalytic converters in automobiles has greatly reduced acid rain producing exhaust products

90. Methyl mercury is a toxin produced indirectly from what energy source?

a. Oil
b. Natural gas
c. Wood
d. Coal

Answers and Explanations

1. B: Na (sodium) is a solid at standard temperature and pressure, which is 0°C (273 K) and 100 kPa (0.986 atm), according to IUPAC. The stronger the intermolecular forces, the greater the likelihood of the material being a solid. Kr and Xe are noble gases and have negligible intermolecular attraction. NH_3 has some hydrogen bonding but is still a gas at STP. Sodium is an alkali metal whose atoms are bonded by metallic bonding and is therefore a solid at STP.

2. D: Generally, the larger and heavier the molecule, the higher the melting point. Decreasing polarity will lower intermolecular attractions and lower the melting point. Long, linear molecules have a larger surface area, and therefore more opportunity to interact with other molecules, which increases the melting point.

3. A: The kinetic energy of the gas molecules is directly proportional to the temperature. If the temperature decreases, so does the molecular motion. A decrease in temperature will not necessarily mean a gas condenses to a liquid. Neither the mass nor the density is impacted, as no material was added or removed, and the volume remained the same.

4. B: The ideal gas law PV=nRT is rearranged to solve for V, and we get V = nRT/P. R is the gas constant, 0.08206 L atm/mol K, and the Celsius temperature must be converted to Kelvin, by adding 273 to 25°C to obtain 298 K. The pressure must be converted to atmospheres, which 101 kPa is essentially 1 atm (0.9967 atm). Plugging the numbers into the equation we get V = 1000 mol (0.08206 L atm/mol K)(298 K)/1 atm, which gives V = 24,453 L. A liter is a cubic decimeter (dm^3) and when converted gives V = 24.5 m^3.

5. D: Since there are twice as many molecules of hydrogen present vs. oxygen, the partial pressure of hydrogen will be greater. The mass of hydrogen will not be greater than the mass of oxygen present even though there are more moles of hydrogen, due to oxygen having a higher molecular weight. Each gas will occupy the same volume. Hydrogen and oxygen gas can coexist in the container without reacting to produce water. There is no indication given that a chemical reaction has occurred.

6. C: Graham's law of diffusion allows one to calculate the relative diffusion rate between two different gases based on their masses.

7. D: London dispersion forces are the weakest intermolecular forces. These interactions occur in all molecules due to unequal electron density around the nucleus, which results in a momentary dipole. Dipole-dipole interactions are those between two polar molecules. The more positive portion of one molecule is attracted to the negative portion of a different molecule. Hydrogen bonding is a stronger type of dipole-dipole interaction which occurs between a hydrogen in one molecule and a nitrogen, oxygen or fluorine atom in another molecule. Hydrogen bonding only occurs between molecules containing H-F, H-O or H-N bonds. Ionic bonds are the strongest intermolecular forces. In ionic molecules, a positive ion is attracted to a negative ion. NaCl is entirely ionic with full charge separation, and the ions are tightly bound to each other in an organized crystalline network.

8. B: Plugging the data into the ideal gas law using the correct units gives the correct answer in atmospheres, which in this case is 2.4 atm. The equation is P = nRT/V. So we have P = 1 mol (0.08206 L atm/mol K)(298 K)/10 L. The R value is 0.08206 L atm/mol K when using L as the volume unit, and delivers the pressure in atm.

9. A: Since there are 7 moles of neon out of a total of 14 moles of gas in the cylinder, the partial pressure of neon will always be 50% of the total pressure, regardless of the temperature.

10. C: To convert from degrees Celsius to Kelvin, add 273. 75° C is equivalent to 348 K. Both X and Y have lower boiling points, which means that they will each boil in the water bath. Z will never become warm enough to boil.

11. D: Both liquids and gases are fluids and therefore flow, but only gases are compressible. The molecules that make up a gas are very far apart, allowing the gas to be compressed into a smaller volume.

12. C: The most electronegative atoms are found near the top right of the periodic table. Fluorine has a high electronegativity, while Francium, located at the bottom left on the table, has a low electrongativity.

13. B: Since both the volume and the temperature remain fixed, the only variable that changes is the number of moles of particles. Because there are now 3 times the number of particles as there were originally, the pressure must increase proportionately and so the pressure must be 3 atm.

14. C: As the temperature drops to -5 °C, the water vapor condenses to a liquid, and then to a solid. The vapor pressure of a solid is much less than that of the corresponding gas. The argon is still a gas at -5 °C, so almost all the pressure in the cylinder is due to argon.

15. A: Heat is absorbed by the solid during melting, therefore ΔH is positive. Going from a solid to a liquid greatly increases the freedom of the particles, therefore increasing the entropy, so ΔS is also positive.

16. B: Freezing is an exothermic event; therefore heat must be given off. The temperature of the material remains unchanged at the freezing point during the process.

17. A: The higher the temperature of the liquid, the greater the solubility of the solid, while the higher the temperature, the lower the solubility of the gas.

18. C: Normality refers to the concentration of acid equivalents (H^+ ions), not the concentration of the solute. 100 g of phosphoric acid has a MW of 98 g/mol. So, 100g/98 g/mol = 1.02 moles of phosphoric acid are in solution. The total volume of the solution is 0.4 L, so the molarity of the solution is 1.02 mol/0.4 L = 2.55 M. Since there are three acid equivalents for every mole of phosphoric acid, the normality is 3 x 2.55 = 7.65 N.

19. D: $AgNO_3$, $NaNO_3$ and NaCl are all highly water soluble and would not precipitate under these conditions. All nitrate compounds and compounds containing Group I metals are soluble in water. AgCl is essentially insoluble in water, and this is the precipitate observed.

20. B: Density is determined by dividing the mass of the solution by its volume. The mass is 200 g, and the total volume is 0.2 L, or 200 mL. So 200 g/200 mL = 1 g/mL.

21. D: 100 mL of a 0.1 M solution of NaOH contains 0.01 moles of NaOH. That means 0.01 moles of acid are required to completely neutralize the solution. The MW of sulfuric acid is 98, so 0.98 g of sulfuric acid is 0.01 mole. But since sulfuric acid has two equivalents of acid per mole, only 0.005 mole of the acid is required or 0.49 g.

22. B: Pure water boils at 100 °C. Water that has salts dissolved in it will boil at a slightly higher temperature, and will conduct electricity much better than pure water.

23. D: Acidity increases as we travel down the periodic table with regard to the halogens. Even though fluorine is the most electronegative element and would be expected to stabilize a negative charge well, it is such a small atom that it is poorly able to stabilize the negative charge and therefore will have a stronger bond to the hydrogen. As the atoms get larger, moving from fluorine to iodine, the ability to stabilize a negative charge becomes greater and the bond with the hydrogen is weaker. A stronger bond with the between the halogen and the hydrogen will result in less acidity, since fewer hydrogen ions will be produced.

24. D: Octane is a nonpolar hydrocarbon with little or no water solubility. Butanol is an alcohol with a small amount of solubility due to its polar –OH group. Ethanol is a smaller, more polar alcohol that is very soluble in water. NaCl is an ionic salt that is highly soluble in water.

25. A: The weight % of the acetic acid is the mass of acetic acid divided by the mass of the acetic acid plus the water. So 50g/(50g +200g) = 0.2, or 20%. The mole fraction is the moles of acetic acid divided by the total number of moles of the solution. So 50 g of acetic acid (MW = 60) is 50g/ 60 g/mol = 0.83 moles. 200 g of water = 11.11 moles. Therefore, 0.83 mol/(0.83 mol + 11.11 mol) = 0.069.

26. C: Since there are three moles of NH_4^+ per mole of salt and 1 mole of PO_4^{3-} per mole of salt, the total ionic concentrations must be 2.7 M of NH_4^+, and 0.9 M of PO_4^{3-}.

27. B: Ethane is an alkane and only very weakly acidic. Methanol, an alcohol, has a slightly acidic proton attached to the oxygen. Acetic acid is much more acidic than methanol with the acidic proton attached to the carboxyl group. Hydrochloric acid is highly acidic and completely dissociates in water.

28. A: Since we have 1 liter of the solution, then 0.02 M represents 0.02 moles of methanol. The mass of methanol can then be found by 0.02mol x MW of CH_3OH (32) = 0.64 g. Molality is the moles of solute (methanol) divided by the number of kilograms of solvent, in this case, it is essentially 1 kg. This is assumed since the solvent is water and the density of water is 1 g/mL. So 0.02 mol/ 1 kg = 0.02 m.

29. D: During osmosis, solvent flows from the lowest to the highest concentration of solute, in this case B to A. The membrane is semi-permeable and only allows the solvent to move, not the solute.

30. B: Beta emission represents the spontaneous decay of a neutron into a proton with the release of an electron. Therefore the resulting nucleus will have one more proton than it did before the reaction, and protons represent the atomic number of an atom. Alpha decay results in the emission of a helium nucleus. The resulting nucleus of an alpha decay would lose two protons and two neutrons, causing a decrease in both the atomic number and the mass number. Gamma decay does not affect the numbers of protons or neutrons in the nucleus. It is an emission of a photon, or packet of energy.

31. C: Since each half life is 2 years, eight years would be 4 half lives. So the mass of material is halved 4 times. Therefore if we start with 1 kg, at two years we would have 0.5 kg, at four years we would have 0.25 kg, after 6 years we would have 0.12 kg, and after 8 years we would have 0.06 kg.

32. D: Using the decay formula, C-14 remaining = C-14 initial$(0.5)^{t/t\ \text{half-life}}$. So, 1 mg $(0.5)^{20000/5730}$ = 0.09 mg. This problem is best solved using the decay formula since 20,000 years is 3.5 half lives. If a student is careful in their reasoning, this problem can be solved without the decay formula. After 3 half-lives, there would be 0.125 mg remaining. If allowed to decay for 4 half-lives, 0.0625 mg would remain. Since only half of this half-life were allowed to elapse, only half of the material would decay, which would be 0.03 mg. Subtracting this amount from 0.125 mg, the amount remaining after 3 half-lives, gives 0.09 mg, which is the amount of material remaining after 3.5 half-lives.

33. D: Isotopes of the same element must have the same chemical behavior, so A, B, and C all represent, in one form or another, chemical behavior. Isotopes differ in mass, and this can be used to separate them by some appropriate physical property.

34. B: Neutrons are neutral in charge, and can impact a nucleus in order to break it.

35. A: Nuclear reactions convert mass into energy ($E = mc^2$). The mass of products is always less than that of the starting materials since some mass is now energy.

36. C: The mass number is the number of protons and the number of neutrons added together. The number of protons is also known as the atomic number and can be found on the periodic table. Therefore, the number of neutrons is the mass number (238) less the number of protons, in this case, 92, so we have 146 neutrons. The number of electrons always equals the number of protons in a neutral atom, so C is the correct answer.

37. D: An alpha particle is a helium nucleus, which contains two protons and two neutrons.

38. B: Phosphorus is in the third period, so the outermost levels must be 3s, 3p. Phosphorus is in Group 5A, which indicates that it has 5 valence electrons. To fill the 3s and 3p, 2 electrons first fill the s orbital, and then the remaining 3 electrons enter the p orbitals. So, $3s^2\ 3p^3$.

39. C: Hund's rule states that electrons must populate empty orbitals of similar energy before pairing up. The Aufbau principle states that electrons must fill lower energy orbitals before filling higher energy orbitals. The Pauli exclusion principle states that no two electrons in the same atom can have the same four quantum numbers, and therefore, two electrons in the same orbital will have opposite spins.

40. D: All of the elements belong to the same row in the periodic table. Atomic radii increase going from right to left in any row of the periodic table. Although these elements belonged to the same row, it is important to also know that atomic radii increase from top to bottom in the groups of the periodic table.

41. A: Sublimation is the process of a solid changing directly into a gas without entering the liquid phase. Fusion refers to a liquid turning into a solid. Diffusion is the process of a material dispersing throughout another. Condensation is generally a gas turning into a liquid.

42. D: Each oxygen has a charge of -2 for a total negative charge of -8. Potassium (K) only exists in compounds as +1. Therefore for the molecule to have a neutral charge, the Mn must be in a +7 oxidation state.

43. B: The trend within any column of the periodic table is that electronegativity decreases going down the column.

44. A: NaCl is an ionic salt, and therefore the most polar. F_2 is nonpolar since the two atoms share the electrons in an equal and symmetrical manner. CH_3OH is an alcohol with a very polar O-H bond. CH_3CH_2Cl is also a polar molecule due to the unequal sharing of electrons between in the C-Cl bond.

45. B: The nitrogen is missing its lone pair of electrons, and should have two dots above it. A correct Lewis structure shows how the atoms are connected to each other as well as all of the valence electrons in the compound. Each bond represents two electrons.

46. C: The more s character the bond has, the shorter it will be. A triple bond is stronger and shorter than a double bond, which is stronger and shorter than a single bond. An sp orbital is found in a triple bond. An sp^2 obital is found in a double bond and sp^3 orbitals are found in single bonds.

47. D: Resonance structures have the same atoms connected to the same atoms, but differ only in electronic structure amongst the atoms. Isomers are molecules that have the same formula but differ in structure. Structural isomers differ in how the atoms are bonded to each other. Stereoisomers are isomers that have the same bonding structure but different arrangements, for example, cis- and trans- isomers.

48. B: Hybrid orbitals arrange themselves to be as far from each other as possible. An sp^2 atom has three hybrid orbitals, so they arrange themselves to be trigonal planar, with 120 ° between the bonds.

49. D: The correct structure of ammonium sulfate is $(NH_4)_2SO_4$. Its molecular weight is 132. The masses of the elements in the compound are: nitrogen 28 (2 x 14), hydrogen 8 (1x8), sulfur 32 (32x1) and oxygen 64 (16x4). To find the percentage composition of each element, divide the element mass by the molecular weight of the compound and multiply by 100. So nitrogen is (28/132)x100 = 21%, hydrogen is (8/132)x100 = 6%, sulfur is (32/132)x100 = 24% and oxygen is (64/132)x100 = 48%.

50. C: Three oxygen are equal to a total charge of -6. Therefore, the two iron atoms must equal that with a positive charge, or +6. So each iron atom must be +3, and the compound is iron (III) oxide.

51. A: By comparing the products to the reactants, there must be at least two Al atoms in the starting material, and at least three sulfate groups. Therefore, a coefficient of 2 must be placed in front of $Al(OH)_3$ and a coefficient of 3 must be placed in front of H_2SO_4. To make the number of hydrogen and oxygen atoms equal on both sides of the equation, a coefficient of 6 must be placed in front of H_2O.

52. D: 1 kg of heptane (MW 100) is equal to 10 moles of heptane. Since 8 moles of water is produced for every mole of heptane reacted, 80 moles of water must be produced. 80 moles of water (MW 18) equals 1440 g, or 1.4 kg.

53. C: 100 g of HBr equals 1.23 moles, and 100 g of Mg equals 4.11 moles. From the coefficients of the balanced equation, the ratio of HBr to Mg is 2:1. This means that to react 1.23 moles of HBr, 2.46 moles of Mg would be required. Since 4.11 moles of Mg are present, Mg is in excess.

54. C: The equation shows that for every liter of methane reacted, one liter of CO_2 and 2 liters of water vapor will be produced. So a total of three liters of gaseous products will be formed for every liter of methane burned. Because the temperature of the reaction products is 200 °C, the water produced will be in vapor (gas) form and not in liquid form. Since 10L of methane were burned, 30 L of gaseous products were formed.

55. B: Since the conversion of B to C is the slow step, this is the only one that determines the reaction rate law. Therefore, the rate law will be based on B, since it is the only reactant in producing C.

56. A: Since the atomic number is 20, which represents the number of protons in the atom, there must be an equal number of electrons in a neutral atom. Protons have a positive charge and electrons are negative. Equal numbers of protons and electrons will result in a neutral atom, or zero charge.

57. C: A pressure increase will force the reaction to go further to the right, which lowers gas pressure to restore equilibrium. Since the water formed is in the liquid phase, it does not appear in the equilibrium equation, so only 1 mole of gas is produced and is part of the equation.

58. D: For a general reaction, $a\,A + b\,B \rightarrow c\,C + d\,D$, the equilibrium equation would take the form:

$$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

where a, b, c and d are the coefficients from the balanced chemical reaction. Pure liquids and solids are excluded from the equation. Since all reactants and products in the problem are gaseous, the equilibrium equation for the reaction would be:

$$K_{eq} = \frac{[CCl_4][HCl]^4}{[CH_4][Cl_2]^4}$$

59. C: Catalysts lower the energy barrier between products and reactants and thus increase the reaction rate.

60. D: 0.5 g of the salt dissolved, which is 0.005 mol of the salt. Since the volume is 1 L, the molarity of the salt is 0.005 M. This means that both species X and Y are present at 0.005 M concentration. The $K_{sp} = [X][Y]$, or [0.005][0.005] which equals 2.5×10^{-5}.

61. A: Lewis acids are compounds capable of accepting a lone pair of electrons. $AlCl_3$ is a very strong Lewis acid and can readily accept a pair of electrons due to Al only having 6 electrons instead of 8 in its outer shell. $FeCl_3$ is also a strong Lewis acid, though milder than $AlCl_3$. Sulfuric acid is a Bronsted-Lowry acid since it produces protons. PCl_3 is a Lewis base since the P can donate its lone pair of electrons to another species.

62. B: NH_3 is ammonia, which is a base. H_2O is amphoteric, meaning that it can act as either a weak acid or a weak base. HF is actually a weak acid, despite fluorine being the most electronegative atom. The small size of the F results in a stronger bond between the H and F, which reduces acidity since this bond will be harder to break. H_3PO_4, phosphoric acid, is high in acidity and HCl is a very strong acid, meaning it completely dissociates.

63. A: First, one must understand that pK_a is the acidity dissociation number. The larger the number, the less acidic. Acetic acid is a carboxylic acid. When H^+ is given off, a negative charge results on the O. Because there is a second equivalent oxygen bonded to the same carbon, this negative charge can be shared between both oxygen atoms. This is known as resonance stabilization and this conjugate base will be more stable and more of the acid molecules will remain dissociated resulting in higher acidity. For ethanol, when the O-H bond breaks, the negative charge resides completely on the O. It cannot be stabilized by other atoms and therefore reforms the methanol rapidly. This results in very low acidity, since very few protons will be released.

64. C: HCl is a strong acid that will completely dissociate. $pH = -\log_{10}[H^+]$, which for this problem is $pH=-\log_{10}(0.1) = 1$. The volume of the solution has no bearing on the pH since we know the concentration.

65. B: The K_a of acetic acid is determined from the pK_a, $K_a = 10^{-pka} = 1.75 \times 10^{-5}$. This is the equilibrium constant for the acetic acid dissociation, or $K_a = [H^+][CH_3COO^-]/[CH_3COOH]$. Using this equilibrium equation to solve for the $[H^+]$, the pH of the buffer can then be found. Solving for the $[H^+]$ concentration, we get $[H^+] = K_a \times [CH_3COOH]/CH_3COO^-]$, or $[H^+] = 1.75 \times 10^{-5} \times [0.1]/[0.2] = 8.75 \times 10^{-6}$. $pH = -\log[H^+] = 5.05$.

66. C: There are 0.05 mol of sulfuric acid being added, but a total of 0.10 mol of H^+ since sulfuric acid is diprotic (H_2SO_4). This is being added to 0.1 mol of NaOH. The moles of acid and base exactly cancel each other out; therefore the pH of the resulting aqueous solution will be near 7.

67. D: To make a buffer, a weak acid and its conjugate base or a weak base and its conjugate acid are commonly used. Buffers work by using the common-ion effect and result in little change in the pH when an acid or a base is added. HCl/NaOH is a strong acid/strong base combination and will not result in a buffer solution. Although the $HNO_3/NaNO_3$ and $H_2SO_4/NaHSO_4$ mixtures are conjugate acid/base pairs, both HNO_3 and H_2SO_4 are strong acids, not weak acids. Neither of these solutions would result in a buffer. Only the NaH_2PO_4/Na_2HPO_4 mixture would result in a buffer as it is a combination of a weak acid and its conjugate base.

68. A: When liquid water changes to steam, the temperature is constant, as in all phase changes. The entropy increases due to the increase in disorder from a liquid to a gas.

69. C: Cooling means heat is leaving the system, so it must be negative. We have 5.9 mol of ammonia cooling 75 °C, or 75 K. So 5.9 mol x -75 K x 35.1 J/(mol)(K) = -15.5 kJ.

70. A: The heat of combustion is determined by subtracting the heats of formation of the reactants from that of the products. So 3(-393.5) + 4(285.8) – (103.8) = -2220.

71. D: In I, dissolving a solid into a liquid breaks up the organized solid matrix, therefore increasing disorder. III converts single particles into two particles, and in IV, solid ice sublimes into a gas. Both of these processes also increase disorder and thus, entropy. II is a decrease in entropy, since 7 molecules, with 3 being gaseous, are reacted to form 2 solid molecules.

72. C: Because all unique materials have differing heat capacities, no two can heat up the same way. All will require different amounts of heat to warm to the same temperature.

73. D: Lead (Pb) goes from a zero oxidation state to a 2+ oxidation state, and is therefore oxidized. Oxidation is the loss of electrons. Hydrogen goes from a 1+ oxidation state to a 0 oxidation state, and is therefore reduced. Reduction is the gaining of electrons.

74. C: Reducing agents give up electrons to another chemical species, which cause that species to gain an electron and become reduced. Oxidizing agents cause another species to be oxidized, or to lose an electron, and are themselves reduced as they gain that electron. Bromine is very electronegative, and is almost always an oxidizing agent. N_2 is nearly inert, or unreactive. Neon is an inert noble gas and would not be a reducing agent. Sodium (Na) is very reactive and eager to give up an electron, and is therefore a good reducing agent in a wide variety of reactions.

75. C: Reduction takes place at the cathode and oxidation takes place at the anode. Mg^{2+} of the salt will be reduced to Mg(0) at the cathode, and Cl^- will be oxidized to Cl_2 at the anode.

76. D: The fact that ΔG for the reaction is negative indicates the reaction is spontaneous. This does not mean the reaction will be faster or slow. Diamonds as we all know do not rapidly convert to graphite, and in fact do so only very slowly, over millions of years, thank goodness.

77. A: The longest straight chain of carbons is four, so the parent name is butane. The alcohol takes number precedence, so it is in the -2- position, placing the methyl in the -3- position. The suffix becomes –ol since it is an alcohol, so the name is 3-methyl-2-butanol.

78. C: The first is an alkyne, which contains a triple bond between carbon atoms. The second is a ketone and contains a carbon-oxygen double bond. The third is an alkene, which has a double bond between two carbon atoms. The fourth is an imide, which contains a double bond between two nitrogen atoms.

79. B: The carbon of a carboxylic acid has three bonds to oxygen atoms and one to a carbon atom. The carbon bonded to the carboxylic carbon will have an oxidation state of zero. Each oxygen atom will have an oxidation number of -2. However, one oxygen is bonded to a hydrogen, which will have an oxidation number of +1. This results in a total oxidation state of -3 for both oxygens bonded to the carbon. Therefore, since the carbon must balance the oxidation states of the oxygens (-3) and the carbon (0), the oxidation state of the carbon must be +3. The three bonds to oxygen give a +3, and the bond to carbon is 0.

80. A: B is 1-butanol, since its longest chain of carbons is 4, not 3. C is 3-pentanone, since there are 5 carbons in the chain and it is a ketone, rather than a carboxylic acid. D is 1-butene, not 3-butene. The name should be assigned by giving the double bond the lowest number.

81. D: Mendeleev was able to connect the trends of the different elements behaviors and develop a table that showed the periodicity of the elements and their relationship to each other.

82. D: Different molecules must have the same chemical formula to be isomers. They differ only in which atoms are bound to which. Having the same molecular weight does not necessarily mean two molecules have the same formula.

83. B: Cis isomers have substituent groups that are on the same side of the molecule across the double bond. Trans isomers are those with substituent groups that are on opposite sides of the molecule across the double bond. I is neither cis nor trans, since both substituents on the same carbon are identical. IV is trans because the two methyl groups are on opposite sides of the molecule. II is cis due to both ethyl groups being on the same side of the molecule. III is also considered cis, although each substituent is different. The heaviest groups on each end of the double bond must be on the same side of the double bond to be cis.

84. A: NMR, or nuclear magnetic resonance, allows one to determine the connectivity of atoms in an organic molecule, by "reading" the resonance signals from the attached hydrogen atoms. IR, or infrared spectroscopy, can help to identify the functional groups that are present, but does not give much information about its position in the molecule. Mass spectrometry breaks apart a large molecule and analyzes the masses of the fragments. It can be useful in analyzing protein structure. HPLC, or high performance liquid chromatography, is a method used to separate a mixture into its components.

85. C: Proteins are large polypeptides, comprised of many amino acids linked together by an amide bond. DNA and RNA are made up of nucleic acids. Carbohydrates are long chains of sugars. Triglycerides are fats and are composed of a glycerol molecule and three fatty acids.

86. C: The closer the data points are to each other, the more precise the data. This does not mean the data is accurate, but that the results are very reproducible.

87. A: Density is mass per volume, typically expressed in units such as g/cm^3, or kg/m^3.

88. B: Acids and bases will react violently if accidentally mixed, as will reducing and oxidizing agents. Both reactions can be highly exothermic and uncontrollable.

89. A: Communities around the world who drink fluoridated water have shown dramatic decreases in the number of dental cavities formed per citizen versus those communities that do not drink fluoridated water.

90. D: Combustion of coal releases significant amounts of Hg into the atmosphere. When the Hg settles into the water, it becomes methylated and concentrates in fish, making them toxic to eat.

Secret Key #1 - Time is Your Greatest Enemy

Pace Yourself

Wear a watch. At the beginning of the test, check the time (or start a chronometer on your watch to count the minutes), and check the time after every few questions to make sure you are "on schedule."

If you are forced to speed up, do it efficiently. Usually one or more answer choices can be eliminated without too much difficulty. Above all, don't panic. Don't speed up and just begin guessing at random choices. By pacing yourself, and continually monitoring your progress against your watch, you will always know exactly how far ahead or behind you are with your available time. If you find that you are one minute behind on the test, don't skip one question without spending any time on it, just to catch back up. Take 15 fewer seconds on the next four questions, and after four questions you'll have caught back up. Once you catch back up, you can continue working each problem at your normal pace.

Furthermore, don't dwell on the problems that you were rushed on. If a problem was taking up too much time and you made a hurried guess, it must be difficult. The difficult questions are the ones you are most likely to miss anyway, so it isn't a big loss. It is better to end with more time than you need than to run out of time.

Lastly, sometimes it is beneficial to slow down if you are constantly getting ahead of time. You are always more likely to catch a careless mistake by working more slowly than quickly, and among very high-scoring test takers (those who are likely to have lots of time left over), careless errors affect the score more than mastery of material.

Secret Key #2 - Guessing is not Guesswork

You probably know that guessing is a good idea. Unlike other standardized tests, there is no penalty for getting a wrong answer. Even if you have no idea about a question, you still have a 20-25% chance of getting it right.

Most test takers do not understand the impact that proper guessing can have on their score. Unless you score extremely high, guessing will significantly contribute to your final score.

Monkeys Take the Test

What most test takers don't realize is that to insure that 20-25% chance, you have to guess randomly. If you put 20 monkeys in a room to take this test, assuming they answered once per question and behaved themselves, on average they would get 20-25% of the questions correct. Put 20 test takers in the room, and the average will be much lower among guessed questions. Why?

1. The test writers intentionally write deceptive answer choices that "look" right. A test taker has no idea about a question, so he picks the "best looking" answer, which is

often wrong. The monkey has no idea what looks good and what doesn't, so it will consistently be right about 20-25% of the time.

2. Test takers will eliminate answer choices from the guessing pool based on a hunch or intuition. Simple but correct answers often get excluded, leaving a 0% chance of being correct. The monkey has no clue, and often gets lucky with the best choice.

This is why the process of elimination endorsed by most test courses is flawed and detrimental to your performance. Test takers don't guess; they make an ignorant stab in the dark that is usually worse than random.

$5 Challenge

Let me introduce one of the most valuable ideas of this course—the $5 challenge:

You only mark your "best guess" if you are willing to bet $5 on it.
You only eliminate choices from guessing if you are willing to bet $5 on it.

Why $5? Five dollars is an amount of money that is small yet not insignificant, and can really add up fast (20 questions could cost you $100). Likewise, each answer choice on one question of the test will have a small impact on your overall score, but it can really add up to a lot of points in the end.

The process of elimination IS valuable. The following shows your chance of guessing it right:

If you eliminate wrong answer choices until only this many remain:	Chance of getting it correct:
1	100%
2	50%
3	33%

However, if you accidentally eliminate the right answer or go on a hunch for an incorrect answer, your chances drop dramatically—to 0%. By guessing among all the answer choices, you are GUARANTEED to have a shot at the right answer.

That's why the $5 test is so valuable. If you give up the advantage and safety of a pure guess, it had better be worth the risk.

What we still haven't covered is how to be sure that whatever guess you make is truly random. Here's the easiest way:

Always pick the first answer choice among those remaining.

Such a technique means that you have decided, **before you see a single test question**, exactly how you are going to guess, and since the order of choices tells you nothing about which one is correct, this guessing technique is perfectly random.

This section is not meant to scare you away from making educated guesses or eliminating choices; you just need to define when a choice is worth eliminating. The $5 test, along with a pre-defined random guessing strategy, is the best way to make sure you reap all of the benefits of guessing.

Secret Key #3 - Practice Smarter, Not Harder

Many test takers delay the test preparation process because they dread the awful amounts of practice time they think necessary to succeed on the test. We have refined an effective method that will take you only a fraction of the time.

There are a number of “obstacles” in the path to success. Among these are answering questions, finishing in time, and mastering test-taking strategies. All must be executed on the day of the test at peak performance, or your score will suffer. The test is a mental marathon that has a large impact on your future.

Just like a marathon runner, it is important to work your way up to the full challenge. So first you just worry about questions, and then time, and finally strategy:

Success Strategy

1. Find a good source for practice tests.
2. If you are willing to make a larger time investment, consider using more than one study guide. Often the different approaches of multiple authors will help you “get” difficult concepts.
3. Take a practice test with no time constraints, with all study helps, “open book.” Take your time with questions and focus on applying strategies.
4. Take a practice test with time constraints, with all guides, “open book.”
5. Take a final practice test without open material and with time limits.

If you have time to take more practice tests, just repeat step 5. By gradually exposing yourself to the full rigors of the test environment, you will condition your mind to the stress of test day and maximize your success.

Secret Key #4 - Prepare, Don’t Procrastinate

Let me state an obvious fact: if you take the test three times, you will probably get three different scores. This is due to the way you feel on test day, the level of preparedness you have, and the version of the test you see. Despite the test writers' claims to the contrary, some versions of the test WILL be easier for you than others.

Since your future depends so much on your score, you should maximize your chances of success. In order to maximize the likelihood of success, you’ve got to prepare in advance. This means taking practice tests and spending time learning the information and test taking strategies you will need to succeed.

Never go take the actual test as a “practice” test, expecting that you can just take it again if you need to. Take all the practice tests you can on your own, but when you go to take the official test, be prepared, be focused, and do your best the first time!

Secret Key #5 - Test Yourself

Everyone knows that time is money. There is no need to spend too much of your time or too little of your time preparing for the test. You should only spend as much of your precious time preparing as is necessary for you to get the score you need.

Once you have taken a practice test under real conditions of time constraints, then you will know if you are ready for the test or not.

If you have scored extremely high the first time that you take the practice test, then there is not much point in spending countless hours studying. You are already there.

Benchmark your abilities by retaking practice tests and seeing how much you have improved. Once you consistently score high enough to guarantee success, then you are ready.

If you have scored well below where you need, then knuckle down and begin studying in earnest. Check your improvement regularly through the use of practice tests under real conditions. Above all, don’t worry, panic, or give up. The key is perseverance!

Then, when you go to take the test, remain confident and remember how well you did on the practice tests. If you can score high enough on a practice test, then you can do the same on the real thing.

General Strategies

The most important thing you can do is to ignore your fears and jump into the test immediately. Do not be overwhelmed by any strange-sounding terms. You have to jump into the test like jumping into a pool—all at once is the easiest way.

Make Predictions

As you read and understand the question, try to guess what the answer will be. Remember that several of the answer choices are wrong, and once you begin reading them, your mind will immediately become cluttered with answer choices designed to throw you off. Your mind is typically the most focused immediately after you have read the question and digested its contents. If you can, try to predict what the correct answer will be. You may be surprised at what you can predict.

Quickly scan the choices and see if your prediction is in the listed answer choices. If it is, then you can be quite confident that you have the right answer. It still won't hurt to check the other answer choices, but most of the time, you've got it!

Answer the Question

It may seem obvious to only pick answer choices that answer the question, but the test writers can create some excellent answer choices that are wrong. Don't pick an answer just because it sounds right, or you believe it to be true. It MUST answer the question. Once you've made your selection, always go back and check it against the question and make sure that you didn't misread the question and that the answer choice does answer the question posed.

Benchmark

After you read the first answer choice, decide if you think it sounds correct or not. If it doesn't, move on to the next answer choice. If it does, mentally mark that answer choice. This doesn't mean that you've definitely selected it as your answer choice, it just means that it's the best you've seen thus far. Go ahead and read the next choice. If the next choice is worse than the one you've already selected, keep going to the next answer choice. If the next choice is better than the choice you've already selected, mentally mark the new answer choice as your best guess.

The first answer choice that you select becomes your standard. Every other answer choice must be benchmarked against that standard. That choice is correct until proven otherwise by another answer choice beating it out. Once you've decided that no other answer choice seems as good, do one final check to ensure that your answer choice answers the question posed.

Valid Information

Don't discount any of the information provided in the question. Every piece of information may be necessary to determine the correct answer. None of the information in the question is there to throw you off (while the answer choices will certainly have information to throw you off). If two seemingly unrelated topics are discussed, don't ignore either. You can be confident there is a relationship, or it wouldn't be included in the question, and you are probably going to have to determine what is that relationship to find the answer.

Avoid "Fact Traps"

Don't get distracted by a choice that is factually true. Your search is for the answer that answers the question. Stay focused and don't fall for an answer that is true but irrelevant. Always go back to the question and make sure you're choosing an answer that actually answers the question and is not just a true statement. An answer can be factually correct, but it MUST answer the question asked. Additionally, two answers can both be seemingly correct, so be sure to read all of the answer choices, and make sure that you get the one that BEST answers the question.

Milk the Question

Some of the questions may throw you completely off. They might deal with a subject you have not been exposed to, or one that you haven't reviewed in years. While your lack of knowledge about the subject will be a hindrance, the question itself can give you many clues that will help you find the correct answer. Read the question carefully and look for clues. Watch particularly for adjectives and nouns describing difficult terms or words that you don't recognize. Regardless of whether you completely understand a word or not, replacing it with a synonym, either provided or one you more familiar with, may help you to understand what the questions are asking. Rather than wracking your mind about specific detailed information concerning a difficult term or word, try to use mental substitutes that are easier to understand.

The Trap of Familiarity

Don't just choose a word because you recognize it. On difficult questions, you may not recognize a number of words in the answer choices. The test writers don't put "make-believe" words on the test, so don't think that just because you only recognize all the words in one answer choice that that answer choice must be correct. If you only recognize words in one answer choice, then focus on that one. Is it correct? Try your best to determine if it is correct. If it is, that's great. If not, eliminate it. Each word and answer choice you eliminate increases your chances of getting the question correct, even if you then have to guess among the unfamiliar choices.

Eliminate Answers

Eliminate choices as soon as you realize they are wrong. But be careful! Make sure you consider all of the possible answer choices. Just because one appears right, doesn't mean that the next one won't be even better! The test writers will usually put more than one good answer choice for every question, so read all of them. Don't worry if you are stuck between two that seem right. By getting down to just two remaining possible choices, your odds are now 50/50. Rather than wasting too much time, play the odds. You are guessing, but guessing wisely because you've been able to knock out some of the answer choices that you know are wrong. If you are eliminating choices and realize that the last answer choice you are left with is also obviously wrong, don't panic. Start over and consider each choice again. There may easily be something that you missed the first time and will realize on the second pass.

Tough Questions

If you are stumped on a problem or it appears too hard or too difficult, don't waste time. Move on! Remember though, if you can quickly check for obviously incorrect answer choices, your chances of guessing correctly are greatly improved. Before you completely give up, at least try to knock out a couple of possible answers. Eliminate what you can and then guess at the remaining answer choices before moving on.

Brainstorm

If you get stuck on a difficult question, spend a few seconds quickly brainstorming. Run through the complete list of possible answer choices. Look at each choice and ask yourself, "Could this answer the question satisfactorily?" Go through each answer choice and consider it independently of the others. By systematically going through all possibilities, you may find something that you would otherwise overlook. Remember though that when you get stuck, it's important to try to keep moving.

Read Carefully

Understand the problem. Read the question and answer choices carefully. Don't miss the question because you misread the terms. You have plenty of time to read each question thoroughly and make sure you understand what is being asked. Yet a happy medium must be attained, so don't waste too much time. You must read carefully, but efficiently.

Face Value

When in doubt, use common sense. Always accept the situation in the problem at face value. Don't read too much into it. These problems will not require you to make huge leaps of logic. The test writers aren't trying to throw you off with a cheap trick. If you have to go beyond creativity and make a leap of logic in order to have an answer choice answer the question, then you should look at the other answer choices. Don't overcomplicate the problem by creating theoretical relationships or explanations that will warp time or space. These are normal problems rooted in reality. It's just that the applicable relationship or explanation may not be readily apparent and you have to figure things out. Use your common sense to interpret anything that isn't clear.

Prefixes

If you're having trouble with a word in the question or answer choices, try dissecting it. Take advantage of every clue that the word might include. Prefixes and suffixes can be a huge help. Usually they allow you to determine a basic meaning. Pre- means before, post-means after, pro - is positive, de- is negative. From these prefixes and suffixes, you can get an idea of the general meaning of the word and try to put it into context. Beware though of any traps. Just because con- is the opposite of pro-, doesn't necessarily mean congress is the opposite of progress!

Hedge Phrases

Watch out for critical hedge phrases, led off with words such as "likely," "may," "can," "sometimes," "often," "almost," "mostly," "usually," "generally," "rarely," and "sometimes." Question writers insert these hedge phrases to cover every possibility. Often an answer choice will be wrong simply because it leaves no room for exception. Unless the situation calls for them, avoid answer choices that have definitive words like "exactly," and "always."

Switchback Words

Stay alert for "switchbacks." These are the words and phrases frequently used to alert you to shifts in thought. The most common switchback word is "but." Others include "although," "however," "nevertheless," "on the other hand," "even though," "while," "in spite of," "despite," and "regardless of."

New Information

Correct answer choices will rarely have completely new information included. Answer choices typically are straightforward reflections of the material asked about and will directly relate to the question. If a new piece of information is included in an answer choice that doesn't even seem to relate to the topic being asked about, then that answer choice is likely incorrect. All of the information needed to answer the question is usually provided for you in the question. You should not have to make guesses that are unsupported or choose answer choices that require unknown information that cannot be reasoned from what is given.

Time Management

On technical questions, don't get lost on the technical terms. Don't spend too much time on any one question. If you don't know what a term means, then odds are you aren't going to get much further since you don't have a dictionary. You should be able to immediately recognize whether or not you know a term. If you don't, work with the other clues that you have—the other answer choices and terms provided—but don't waste too much time trying to figure out a difficult term that you don't know.

Contextual Clues

Look for contextual clues. An answer can be right but not the correct answer. The contextual clues will help you find the answer that is most right and is correct. Understand the context in which a phrase or statement is made. This will help you make important distinctions.

Don't Panic

Panicking will not answer any questions for you; therefore, it isn't helpful. When you first see the question, if your mind goes blank, take a deep breath. Force yourself to mechanically go through the steps of solving the problem using the strategies you've learned.

Pace Yourself

Don't get clock fever. It's easy to be overwhelmed when you're looking at a page full of questions, your mind is full of random thoughts and feeling confused, and the clock is ticking down faster than you would like. Calm down and maintain the pace that you have set for yourself. As long as you are on track by monitoring your pace, you are guaranteed to have enough time for yourself. When you get to the last few minutes of the test, it may seem like you won't have enough time left, but if you only have as many questions as you should have left at that point, then you're right on track!

Answer Selection

The best way to pick an answer choice is to eliminate all of those that are wrong, until only one is left and confirm that is the correct answer. Sometimes though, an answer choice may immediately look right. Be careful! Take a second to make sure that the other choices are not equally obvious. Don't make a hasty mistake. There are only two times that you should stop before checking other answers. First is when you are positive that the answer choice you have selected is correct. Second is when time is almost out and you have to make a quick guess!

Check Your Work

Since you will probably not know every term listed and the answer to every question, it is important that you get credit for the ones that you do know. Don't miss any questions through careless mistakes. If at all possible, try to take a second to look back over your answer selection and make sure you've selected the correct answer choice and haven't made a costly careless mistake (such as marking an answer choice that you didn't mean to mark). The time it takes for this quick double check should more than pay for itself in caught mistakes.

Beware of Directly Quoted Answers

Sometimes an answer choice will repeat word for word a portion of the question or reference section. However, beware of such exact duplication. It may be a trap! More than likely, the correct choice will paraphrase or summarize a point, rather than being exactly the same wording.

Slang

Scientific sounding answers are better than slang ones. An answer choice that begins “To compare the outcomes…” is much more likely to be correct than one that begins “Because some people insisted…”

Extreme Statements

Avoid wild answers that throw out highly controversial ideas that are proclaimed as established fact. An answer choice that states the “process should used in certain situations, if…” is much more likely to be correct than one that states the “process should be discontinued completely.” The first is a calm rational statement and doesn’t even make a definitive, uncompromising stance, using a hedge word “if” to provide wiggle room, whereas the second choice is a radical idea and far more extreme.

Answer Choice Families

When you have two or more answer choices that are direct opposites or parallels, one of them is usually the correct answer. For instance, if one answer choice states “x increases” and another answer choice states “x decreases” or “y increases,” then those two or three answer choices are very similar in construction and fall into the same family of answer choices. A family of answer choices consists of two or three answer choices, very similar in construction, but often with directly opposite meanings. Usually the correct answer choice will be in that family of answer choices. The “odd man out” or answer choice that doesn’t seem to fit the parallel construction of the other answer choices is more likely to be incorrect.

Special Report: What is Test Anxiety and How to Overcome It?

The very nature of tests caters to some level of anxiety, nervousness, or tension, just as we feel for any important event that occurs in our lives. A little bit of anxiety or nervousness can be a good thing. It helps us with motivation, and makes achievement just that much sweeter. However, too much anxiety can be a problem, especially if it hinders our ability to function and perform.

"Test anxiety," is the term that refers to the emotional reactions that some test-takers experience when faced with a test or exam. Having a fear of testing and exams is based upon a rational fear, since the test-taker's performance can shape the course of an academic career. Nevertheless, experiencing excessive fear of examinations will only interfere with the test-taker's ability to perform and chance to be successful.

There are a large variety of causes that can contribute to the development and sensation of test anxiety. These include, but are not limited to, lack of preparation and worrying about issues surrounding the test.

Lack of Preparation

Lack of preparation can be identified by the following behaviors or situations:

Not scheduling enough time to study, and therefore cramming the night before the test or exam
Managing time poorly, to create the sensation that there is not enough time to do everything
Failing to organize the text information in advance, so that the study material consists of the entire text and not simply the pertinent information
Poor overall studying habits

Worrying, on the other hand, can be related to both the test taker, or many other factors around him/her that will be affected by the results of the test. These include worrying about:

Previous performances on similar exams, or exams in general
How friends and other students are achieving
The negative consequences that will result from a poor grade or failure

There are three primary elements to test anxiety. Physical components, which involve the same typical bodily reactions as those to acute anxiety (to be discussed below). Emotional factors have to do with fear or panic. Mental or cognitive issues concerning attention spans and memory abilities.

Physical Signals

There are many different symptoms of test anxiety, and these are not limited to mental and emotional strain. Frequently there are a range of physical signals that will let a test taker know that he/she is suffering from test anxiety. These bodily changes can include the following:

Perspiring
Sweaty palms
Wet, trembling hands
Nausea
Dry mouth
A knot in the stomach
Headache
Faintness
Muscle tension
Aching shoulders, back and neck
Rapid heart beat
Feeling too hot/cold

To recognize the sensation of test anxiety, a test-taker should monitor him/herself for the following sensations:

The physical distress symptoms as listed above
Emotional sensitivity, expressing emotional feelings such as the need to cry or laugh too much, or a sensation of anger or helplessness
A decreased ability to think, causing the test-taker to blank out or have racing thoughts that are hard to organize or control.

Though most students will feel some level of anxiety when faced with a test or exam, the majority can cope with that anxiety and maintain it at a manageable level. However, those who cannot are faced with a very real and very serious condition, which can and should be controlled for the immeasurable benefit of this sufferer.

Naturally, these sensations lead to negative results for the testing experience. The most common effects of test anxiety have to do with nervousness and mental blocking.

Nervousness

Nervousness can appear in several different levels:

The test-taker's difficulty, or even inability to read and understand the questions on the test
The difficulty or inability to organize thoughts to a coherent form
The difficulty or inability to recall key words and concepts relating to the testing questions (especially essays)
The receipt of poor grades on a test, though the test material was well known by the test taker

Conversely, a person may also experience mental blocking, which involves:

Blanking out on test questions
Only remembering the correct answers to the questions when the test has already finished.

Fortunately for test anxiety sufferers, beating these feelings, to a large degree, has to do with proper preparation. When a test taker has a feeling of preparedness, then anxiety will be dramatically lessened.

The first step to resolving anxiety issues is to distinguish which of the two types of anxiety are being suffered. If the anxiety is a direct result of a lack of preparation, this should be considered a normal reaction, and the anxiety level (as opposed to the test results) shouldn't be anything to worry about. However, if, when adequately prepared, the test-taker still panics, blanks out, or seems to overreact, this is not a fully rational reaction. While this can be considered normal too, there are many ways to combat and overcome these effects.

Remember that anxiety cannot be entirely eliminated, however, there are ways to minimize it, to make the anxiety easier to manage. Preparation is one of the best ways to minimize test anxiety. Therefore the following techniques are wise in order to best fight off any anxiety that may want to build.

To begin with, try to avoid cramming before a test, whenever it is possible. By trying to memorize an entire term's worth of information in one day, you'll be shocking your system, and not giving yourself a very good chance to absorb the information. This is an easy path to anxiety, so for those who suffer from test anxiety, cramming should not even be considered an option.

Instead of cramming, work throughout the semester to combine all of the material which is presented throughout the semester, and work on it gradually as the course goes by, making sure to master the main concepts first, leaving minor details for a week or so before the test.

To study for the upcoming exam, be sure to pose questions that may be on the examination, to gauge the ability to answer them by integrating the ideas from your texts, notes and lectures, as well as any supplementary readings.

If it is truly impossible to cover all of the information that was covered in that particular term, concentrate on the most important portions, that can be covered very well. Learn these concepts as best as possible, so that when the test comes, a goal can be made to use these concepts as presentations of your knowledge.

In addition to study habits, changes in attitude are critical to beating a struggle with test anxiety. In fact, an improvement of the perspective over the entire test-taking experience can actually help a test taker to enjoy studying and therefore improve the overall experience. Be certain not to overemphasize the significance of the grade - know that the result of the test is neither a reflection of self worth, nor is it a measure of intelligence; one grade will not predict a person's future success.

To improve an overall testing outlook, the following steps should be tried:

Keeping in mind that the most reasonable expectation for taking a test is to expect to try to demonstrate as much of what you know as you possibly can.

Reminding ourselves that a test is only one test; this is not the only one, and there will be others.

The thought of thinking of oneself in an irrational, all-or-nothing term should be avoided at all costs.

A reward should be designated for after the test, so there's something to look forward to. Whether it be going to a movie, going out to eat, or simply visiting friends, schedule it in advance, and do it no matter what result is expected on the exam.

Test-takers should also keep in mind that the basics are some of the most important things, even beyond anti-anxiety techniques and studying. Never neglect the basic social, emotional and biological needs, in order to try to absorb information. In order to best achieve, these three factors must be held as just as important as the studying itself.

Study Steps

Remember the following important steps for studying:

Maintain healthy nutrition and exercise habits. Continue both your recreational activities and social pass times. These both contribute to your physical and emotional well being.
Be certain to get a good amount of sleep, especially the night before the test, because when you're overtired you are not able to perform to the best of your best ability.
Keep the studying pace to a moderate level by taking breaks when they are needed, and varying the work whenever possible, to keep the mind fresh instead of getting bored.
When enough studying has been done that all the material that can be learned has been learned, and the test taker is prepared for the test, stop studying and do something relaxing such as listening to music, watching a movie, or taking a warm bubble bath.

There are also many other techniques to minimize the uneasiness or apprehension that is experienced along with test anxiety before, during, or even after the examination. In fact, there are a great deal of things that can be done to stop anxiety from interfering with lifestyle and performance. Again, remember that anxiety will not be eliminated entirely, and it shouldn't be. Otherwise that "up" feeling for exams would not exist, and most of us depend on that sensation to perform better than usual. However, this anxiety has to be at a level that is manageable.

Of course, as we have just discussed, being prepared for the exam is half the battle right away. Attending all classes, finding out what knowledge will be expected on the exam, and knowing the exam schedules are easy steps to lowering anxiety. Keeping up with work will remove the need to cram, and efficient study habits will eliminate wasted time. Studying should be done in an ideal location for concentration, so that it is simple to become interested in the material and give it complete attention. A method such as

SQ3R (Survey, Question, Read, Recite, Review) is a wonderful key to follow to make sure that the study habits are as effective as possible, especially in the case of learning from a textbook. Flashcards are great techniques for memorization. Learning to take good notes will mean that notes will be full of useful information, so that less sifting will need to be done to seek out what is pertinent for studying. Reviewing notes after class and then again on occasion will keep the information fresh in the mind. From notes that have been taken summary sheets and outlines can be made for simpler reviewing.

A study group can also be a very motivational and helpful place to study, as there will be a sharing of ideas, all of the minds can work together, to make sure that everyone understands, and the studying will be made more interesting because it will be a social occasion.

Basically, though, as long as the test-taker remains organized and self confident, with efficient study habits, less time will need to be spent studying, and higher grades will be achieved.

To become self confident, there are many useful steps. The first of these is "self talk." It has been shown through extensive research, that self-talk for students who suffer from test anxiety, should be well monitored, in order to make sure that it contributes to self confidence as opposed to sinking the student. Frequently the self talk of test-anxious students is negative or self-defeating, thinking that everyone else is smarter and faster, that they always mess up, and that if they don't do well, they'll fail the entire course. It is important to decreasing anxiety that awareness is made of self talk. Try writing any negative self thoughts and then disputing them with a positive statement instead. Begin self-encouragement as though it was a friend speaking. Repeat positive statements to help reprogram the mind to believing in successes instead of failures.

Helpful Techniques

Other extremely helpful techniques include:

Self-visualization of doing well and reaching goals

While aiming for an "A" level of understanding, don't try to "overprotect" by setting your expectations lower. This will only convince the mind to stop studying in order to meet the lower expectations.

Don't make comparisons with the results or habits of other students. These are individual factors, and different things work for different people, causing different results.

Strive to become an expert in learning what works well, and what can be done in order to improve. Consider collecting this data in a journal.

Create rewards for after studying instead of doing things before studying that will only turn into avoidance behaviors.

Make a practice of relaxing - by using methods such as progressive relaxation, self-hypnosis, guided imagery, etc - in order to make relaxation an automatic sensation. Work on creating a state of relaxed concentration so that concentrating will take on the focus of the mind, so that none will be wasted on worrying.

Take good care of the physical self by eating well and getting enough sleep.

Plan in time for exercise and stick to this plan.

Beyond these techniques, there are other methods to be used before, during and after the test that will help the test-taker perform well in addition to overcoming anxiety.

Before the exam comes the academic preparation. This involves establishing a study schedule and beginning at least one week before the actual date of the test. By doing this, the anxiety of not having enough time to study for the test will be automatically eliminated. Moreover, this will make the studying a much more effective experience, ensuring that the learning will be an easier process. This relieves much undue pressure on the test-taker.

Summary sheets, note cards, and flash cards with the main concepts and examples of these main concepts should be prepared in advance of the actual studying time. A topic should never be eliminated from this process. By omitting a topic because it isn't expected to be on the test is only setting up the test-taker for anxiety should it actually appear on the exam. Utilize the course syllabus for laying out the topics that should be studied. Carefully go over the notes that were made in class, paying special attention to any of the issues that the professor took special care to emphasize while lecturing in class. In the textbooks, use the chapter review, or if possible, the chapter tests, to begin your review.

It may even be possible to ask the instructor what information will be covered on the exam, or what the format of the exam will be (for example, multiple choice, essay, free form, true-false). Additionally, see if it is possible to find out how many questions will be on the test. If a review sheet or sample test has been offered by the professor, make good use of it, above anything else, for the preparation for the test. Another great resource for getting to know the examination is reviewing tests from previous semesters. Use these tests to review, and aim to achieve a 100% score on each of the possible topics. With a few exceptions, the goal that you set for yourself is the highest one that you will reach.

Take all of the questions that were assigned as homework, and rework them to any other possible course material. The more problems reworked, the more skill and confidence will form as a result. When forming the solution to a problem, write out each of the steps. Don't simply do head work. By doing as many steps on paper as possible, much clarification and therefore confidence will be formed. Do this with as many homework problems as possible, before checking the answers. By checking the answer after each problem, a reinforcement will exist, that will not be on the exam. Study situations should be as exam-like as possible, to prime the test-taker's system for the experience. By waiting to check the answers at the end, a psychological advantage will be formed, to decrease the stress factor.

Another fantastic reason for not cramming is the avoidance of confusion in concepts, especially when it comes to mathematics. 8-10 hours of study will become one hundred percent more effective if it is spread out over a week or at least several days, instead of doing it all in one sitting. Recognize that the human brain requires time in order to assimilate new material, so frequent breaks and a span of study time over several days will be much more beneficial.

Additionally, don't study right up until the point of the exam. Studying should stop a minimum of one hour before the exam begins. This allows the brain to rest and put things in their proper order. This will also provide the time to become as relaxed as possible when going into the examination room. The test-taker will also have time to eat well and eat sensibly. Know that the brain needs food as much as the rest of the body. With enough food and enough sleep, as well as a relaxed attitude, the body and the mind are primed for success.

Avoid any anxious classmates who are talking about the exam. These students only spread anxiety, and are not worth sharing the anxious sentimentalities.

Before the test also involves creating a positive attitude, so mental preparation should also be a point of concentration. There are many keys to creating a positive attitude. Should fears become rushing in, make a visualization of taking the exam, doing well, and seeing an A written on the paper. Write out a list of affirmations that will bring a feeling of confidence, such as "I am doing well in my English class," "I studied well and know my material," "I enjoy this class." Even if the affirmations aren't believed at first, it sends a positive message to the subconscious which will result in an alteration of the overall belief system, which is the system that creates reality.

If a sensation of panic begins, work with the fear and imagine the very worst! Work through the entire scenario of not passing the test, failing the entire course, and dropping out of school, followed by not getting a job, and pushing a shopping cart through the dark alley where you'll live. This will place things into perspective! Then, practice deep breathing and create a visualization of the opposite situation - achieving an "A" on the exam, passing the entire course, receiving the degree at a graduation ceremony.

On the day of the test, there are many things to be done to ensure the best results, as well as the most calm outlook. The following stages are suggested in order to maximize test-taking potential:

Begin the examination day with a moderate breakfast, and avoid any coffee or beverages with caffeine if the test taker is prone to jitters. Even people who are used to managing caffeine can feel jittery or light-headed when it is taken on a test day.

Attempt to do something that is relaxing before the examination begins. As last minute cramming clouds the mastering of overall concepts, it is better to use this time to create a calming outlook.

Be certain to arrive at the test location well in advance, in order to provide time to select a location that is away from doors, windows and other distractions, as well as giving enough time to relax before the test begins.

Keep away from anxiety generating classmates who will upset the sensation of stability and relaxation that is being attempted before the exam.

Should the waiting period before the exam begins cause anxiety, create a self-distraction by reading a light magazine or something else that is relaxing and simple.

During the exam itself, read the entire exam from beginning to end, and find out how much time should be allotted to each individual problem. Once writing the exam, should more time be taken for a problem, it should be abandoned, in order to begin another problem. If there is time at the end, the unfinished problem can always be returned to and completed.

Read the instructions very carefully - twice - so that unpleasant surprises won't follow during or after the exam has ended.

When writing the exam, pretend that the situation is actually simply the completion of homework within a library, or at home. This will assist in forming a relaxed atmosphere, and will allow the brain extra focus for the complex thinking function.

Begin the exam with all of the questions with which the most confidence is felt. This will build the confidence level regarding the entire exam and will begin a quality momentum. This will also create encouragement for trying the problems where uncertainty resides.

Going with the "gut instinct" is always the way to go when solving a problem. Second guessing should be avoided at all costs. Have confidence in the ability to do well.

For essay questions, create an outline in advance that will keep the mind organized and make certain that all of the points are remembered. For multiple choice, read every answer, even if the correct one has been spotted - a better one may exist.

Continue at a pace that is reasonable and not rushed, in order to be able to work carefully. Provide enough time to go over the answers at the end, to check for small errors that can be corrected.

Should a feeling of panic begin, breathe deeply, and think of the feeling of the body releasing sand through its pores. Visualize a calm, peaceful place, and include all of the sights, sounds and sensations of this image. Continue the deep breathing, and take a few minutes to continue this with closed eyes. When all is well again, return to the test.

If a "blanking" occurs for a certain question, skip it and move on to the next question. There will be time to return to the other question later. Get everything done that can be done, first, to guarantee all the grades that can be compiled, and to build all of the confidence possible. Then return to the weaker questions to build the marks from there.

Remember, one's own reality can be created, so as long as the belief is there, success will follow. And remember: anxiety can happen later, right now, there's an exam to be written!

After the examination is complete, whether there is a feeling for a good grade or a bad grade, don't dwell on the exam, and be certain to follow through on the reward that was promised…and enjoy it! Don't dwell on any mistakes that have been made, as there is nothing that can be done at this point anyway.

Additionally, don't begin to study for the next test right away. Do something relaxing for a while, and let the mind relax and prepare itself to begin absorbing information again.

From the results of the exam - both the grade and the entire experience, be certain to learn from what has gone on. Perfect studying habits and work some more on confidence in order to make the next examination experience even better than the last one.

Learn to avoid places where openings occurred for laziness, procrastination and day dreaming.

Use the time between this exam and the next one to better learn to relax, even learning to relax on cue, so that any anxiety can be controlled during the next exam. Learn how to relax the body. Slouch in your chair if that helps. Tighten and then relax all of the different muscle groups, one group at a time, beginning with the feet and then working all the way up to the neck and face. This will ultimately relax the muscles more than they were to begin with. Learn how to breathe deeply and comfortably, and focus on this breathing going in and out as a relaxing thought. With every exhale, repeat the word "relax."

As common as test anxiety is, it is very possible to overcome it. Make yourself one of the test-takers who overcome this frustrating hindrance.

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