Chapter 1: Introduction to Chemistry

Matter: Anything that occupies space and has mass.

Mass: The quantity of matter which a substance possesses and, depending on the gravitational force acting on it, has a unit of weight assigned to it.

Since matter does occupy space, we can compare the masses of various substances that occupy a particular unit volume. This relationship of mass to a unit volume is called the **density** of the substance. It can be expressed in a mathematical formula as $D = \frac{m}{v}$. The basic unit of mass (m) in chemistry is the gram (g), and of volume (V) is the cubic centimeter (cm^3) or milliliter (mL).

States of Matter

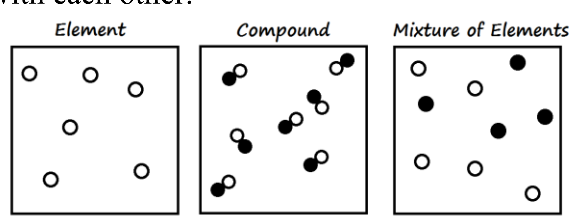
Matter occurs in three states: solid, liquid, and gas. These states of matter can often be changed by the addition or subtraction of heat energy.

- A **solid** has both a definite size and a definite shape.
- A **liquid** has a definite volume but takes the shape of the container.
- A **gas** has neither a definite shape nor a definite volume.

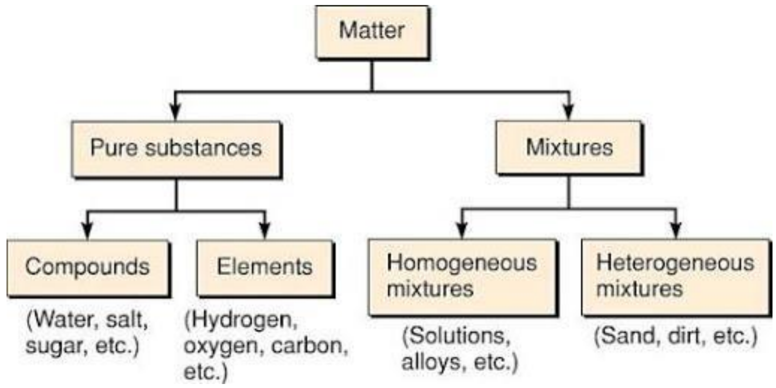
Composition of Matter

Matter can be divided into 2 general categories:

- A **distinct/pure substance** can be subdivided into the smallest particle that still has the properties of that substance, it has only one type of particle alone or bonded as a group
    - If the substance is made up of only one kind of atom and cannot be broken down further, it is called an **element**.
        - Atoms are considered to be the basic building blocks of matter that can't be easily created or destroyed.
        - Diatomic elements are 2 atom elements: hydrogen, nitrogen, oxygen, fluorine, chlorine, iodine, bromine
    - If two or more kinds of atoms are chemically combined in a fixed proportion, it is called a **compound**.
        - The smallest naturally occurring unit of a compound is called a molecule. A molecule of a compound has a definite shape that is determined by how the atoms are bonded to or combined with each other.



- Mixtures, however, can vary in their composition. A mixture is a physical combination of 2 or more substances in which proportions may vary. Mixtures can also be separated into 2 categories:
    - Homogenous: Particles are distributed evenly throughout and there are no visible differences/looks the same throughout
    - Heterogenous: Particles are not distributed evenly and do not look the same



Physical Separation:

- Evaporation: used to contain a soluble (dissolved) solid from a liquid (water)
- Filtration: separate out solid form from a solution/liquid, separation based on particle size not separating homogeneous solutions
- Distillation: used to separate two liquids, collects liquid (based on boiling point)
- Chromatography: separation based on solubility (how well it dissolves)

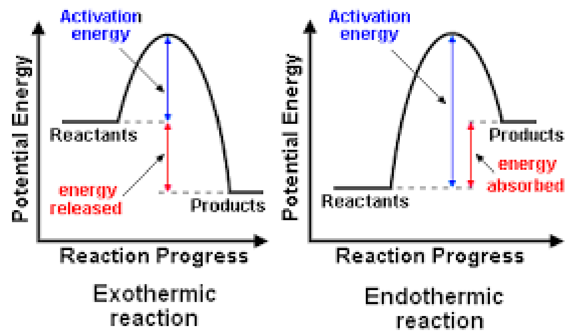
Chemical and Physical Properties

- **Physical properties** of matter are those properties that can usually be observed with our senses. They include everything about a substance that can be noted when no change is occurring in the type of structure that makes up its smallest component.
- Some common examples are physical state, color, odor, solubility in water, density, melting point, taste, boiling point, and hardness.
  Ex: boiling point, freezing point, density, color, volume, mass
- **Chemical properties** are those properties that can be observed in regard to whether or not a substance changes chemically, often as a result of reacting with other substances.
- Some common examples are: iron rusts in moist air, nitrogen does not burn, gold does not rust, sodium reacts with water, silver does not react with water, and water can be decomposed by an electric current.
  Ex: Iron rusts, paper burns

Chemical and Physical Changes

- The changes matter undergoes are classified as either physical or chemical.
- In general, a **physical change** alters some aspect of the physical properties of matter, but the composition remains constant. The most often altered properties are form and state.
- Some examples of physical changes are breaking class, cutting wood, melting ice, and magnetizing a piece of metal.
- In some cases, the process that caused the change can be easily reversed and the substance regains its original form.
- Water changing its state is a good example of physical changes. In the solid state, ice, water has a definite shape and size. As heat is added, it changes to the liquid state, where it has a definite volume but takes the shape of the container. When water is heated above its boiling point, it changes to steam. Steam, a gas, has neither a definite size, because it fills the containing space, nor shape, because it takes the shape of the container.
  Ex: $2\ H_2O \rightarrow 2\ H_2 + O_2$

- **Chemical changes** are changes in the composition and structure of a substance. They are always accompanied by energy changes.
  Ex: solid ↔ liquid ↔ gas or H_2O (solid) → H_2O (gas)
- If the energy released in the formation of a new substance exceeds the chemical energy in the original substances, energy will be given off, usually in the form of heat or light or both. This is called an **exothermic reaction**. If, however, the new structure needs to absorb more energy than is available from the reactants, the result is an **endothermic reaction**.



Conservation of Mass

- When ordinary chemical changes occur, the mass of the reactants equals the mass of the products. This can be stated in another way: In a chemical change, matter can neither be created nor destroyed, but only changed from one form to another (the **Law of Conservation of Matter**).

Energy

Definition of Energy

- Energy is defined as the capacity to do work. Work is done whenever a force is applied over a distance.
- Therefore, anything that can force matter to move, to change speed, or to change direction has energy. Work itself is measured in joules, and so is energy.
- In some problems, however, energy may be expressed in kilocalories. The relationship between these two units is 4.18×10^8 joules (J) equals 1 kilocalorie (kcal).

Forms of Energy

- Energy may appear in a variety of forms.
- Most commonly, energy in reactions is evolved as heat.
- Some other forms of energy are light, sound, mechanical energy, electrical energy, and chemical energy.
- Energy can be converted from one form to another.
- Two general classifications of energy are potential energy and kinetic energy. **Potential energy** is stored energy due to overwhelming forces in nature. **Kinetic energy** is energy of motion.

Types of Reactions

- When physical or chemical changes occur, energy changes are involved.
- Change of heat content can be designed as ΔH. The heat content (H) is sometimes referred to as enthalpy.

- Every system has a certain amount of heat. The changes during the course of a physical or chemical change. The change in heat content, ΔH, is the difference between the heat content of the products and that of the reactants. The equation is

$$\boxed{\Delta H_{rxn} = H_{products} - H_{reactants}}$$

- If the heat content of the products is greater than the heat content of the reactants, ΔH is a positive quantity ($\Delta H > 0$) and the reaction is endothermic.
- If, however, the heat content of the products is less than the heat content of the reactants, ΔH is a negative quantity ($\Delta H < 0$) and the reaction is exothermic.

Conservation of Energy

- Experiments have shown that energy is neither gained nor lost in physical or chemical changes. This principle is known as the Law of Conservation of Energy and is often stated as follows: Energy is neither created nor destroyed in ordinary physical or chemical changes. If the system under study loses energy, the reaction is exothermic and the ΔH is negative.
- Therefore, the system's surroundings must gain the energy that the system loses so that energy is conserved.

Measurements and Calculations

Metric System

It is important that scientists around the world use the same units when communicating information. For this reason, scientists use the modernized metric system, commonly known as **SI**.

Temperature Measurements

The most commonly used temperature scale in scientific work is the Celsius scale. Another scale is based on the lowest theoretical temperature (called absolute zero) referred to the Kelvin scale.

K = °C +273°   °C = K - 273°

Heat Measurements

- **Heat energy** (or just heat) is a form of energy that transfers among particles in a substance (or system) by means of the kinetic energy of those particles. In other words, heat is transferred by particles bouncing into each other.
- A **calorie** is used to measure the quantity of heat and is defined as the amount of heat needed to raise the temperature of 1 gram of water by 1 degree on the Celsius scale. This is a rather small unit to measure the quantity of heat involved in most chemical reactions.
- Therefore, the **kilocalorie** is more often used. The kilocalorie equals 1,000 calories. It is the quantity of heat that will increase the temperature of 1 kilogram of water by 1 degree on the Celsius scale.
- Although the calorie is commonly used in everyday usage with regard to food, the SI unit for heat energy is the **joule**. It is abbreviated as J and, because it is a rather small unit, it is commonly given in kilojoules (kJ). This relationship between the calorie and the joule is that 1 calorie equals 4.18 joules.

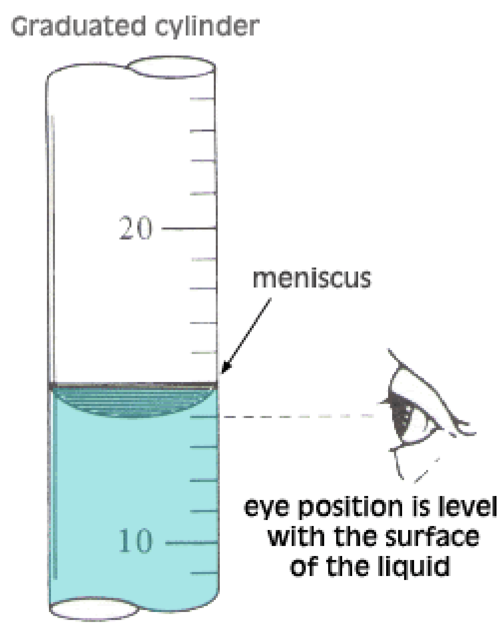
Dimensional Analysis (Factor-Label Method of Conversion)

- When you are working problems that involve numbers with units of measurement, it is convenient to use this method so that you do not become confused in the operations of multiplication or division.
- For example, if you are changing 0.001 kilogram to milligrams, you set up each conversion as a fraction so that all units will factor out except the one you want in the answer.

$$1 \times 10^{-3}\ kg \times \frac{1 \times 10^{3} g}{1\ kg} \times \frac{1 \times 10^{3}\ mg}{1\ g} = 1 \times 10^{3} mg$$

- Notice that the kilogram is made the denominator in the first fraction to be factored with the original kilogram unit. The numerator is equal to the denominator except that the numerator is expressed in smaller units. The second fraction has the gram unit in the denominator to be factored with the gram unit in the preceding fraction. The answer is in milligrams because this is the only unit remaining and it assures you that the correct operations have been performed in the conversation.
  1 ft = ? cm
  $$1\ ft \times \frac{12\ in,}{1\ ft} \times \frac{2.54\ cm}{1\ in} = 30.48\ cm$$

Precision, Accuracy, and Uncertainty



- Two other factors to consider in measurement are precision and accuracy.
- **Precision** indicates the reliability or reproducibility of a measurement.
- **Accuracy** indicates how close a measurement is to its known or accepted value.
- Regardless of precision and accuracy, all measurements have a degree of uncertainty.
- This is usually dependent on one or both of two factors - the limitation of the measuring instrument and the skill of the person making the measurement.
- The graduated cylinder in the illustration contains a quantity of water to be measured. It is obvious that the quantity is between 10 and 20 millimeters because the meniscus lies between these two marked quantities.
- Now, checking to see where the bottom of the meniscus lies within reference to the ten intervening subdivisions, we see that it is between the fourth and fifth. This means that the volume lies between 14 and 15 millimeters.

Significant Figures

- Any time a measurement is recorded, it includes all the digits that are certain plus one uncertain digit. These certain digits plus the one uncertain digit are referred to as **significant figures**.
- The more digits you are able to record in a measurement, the less relative uncertainty there is in the measurement. The following table summarizes the rules of significant figures.

Rule	Example	Number of Significant Figures

All digits other than zeros are significant.	25 g 5.471 g	2 4
Zeros between nonzero digits are significant.	309 g 40.06 g	3 4
Final zeros to the right of the decimal point are significant.	6.00 mL 2.350 mL	3 4
In numbers smaller than 1, zeros to the left or directly to the right of the decimal point are not significant.	0.05 cm 0.060 cm	1. The zeros merely mark the position of the decimal point. 2. The first two zeros mark the position of the decimal point. The final zero is significant.

- One last rule deals with final zeros in a whole number. These zeros may or may not be significant, depending on the measuring instrument.
- For instance, if an instrument that measures to the nearest mile is used, the number 3,000 miles has four significant figures.
- If, however, the instrument in question records miles to the nearest thousands, there is only one significant figure.
- The number of significant figures in 3,000 could be one, two, three, or four, depending on the limitation of the measuring device.

This problem can be avoided by using the system of scientific notation. For this example, the following notations would rather indicate the numbers of significant figures:

3×10^3 one significant figure
3.0×10^3 two significant figures
3.00×10^3 three significant figures
3.000×10^3 four significant figures

Calculations with Significant Figures

- When you do calculations involving numbers that do not have the same number of significant figures in each, keep the following two rules in mind.
- First, in multiplication and division, the number of significant figures in a product or quotient of measured quantities is the same as the number of significant figures in the quantity having the smaller number of significant figures.

EXAMPLE 1:

Problem	Unrounded Answer	Answer rounded to the correct number of sig figs
4.29 cm ×3.24 cm =	13.8996 cm^2=	13.9 cm^2

Explanation: Both measured quantities have three significant figures. Therefore, the answer should be rounded to three significant figures.

EXAMPLE 2:

Problem	Unrounded Answer	Answer rounded to the correct number of sig figs
4.29 cm ×3.2 cm =	13.728 cm^2=	14 cm^2

Explanation: One of the measured quantities has only two significant figures. Therefore, the answer should be rounded to two significant figures.

EXAMPLE 3:

Problem	Unrounded Answer	Answer rounded to the correct number of sig figs
8.47 cm^2/ 4.26 cm =	1.9882629 cm =	1.99 cm

Explanation: Both measured quantities have three significant figures. Therefore, the answer should be rounded to three significant figures.

Second, when adding or subtracting measured quantities, the sum or difference should be rounded to the same number of decimal places as the quantity having the least number of decimal places.

EXAMPLE 1:

Problem	Unrounded Answer	Answer rounded to the correct number of sig figs
$3.56\ cm\ +\ 2.6\ cm\ +\ 6.12\ cm$ =	12.28 cm =	12.3 cm

Explanation: One of the quantities added has only one decimal place. Therefore, the answer should be rounded to only one decimal place.

EXAMPLE 2:

Problem	Unrounded Answer	Answer rounded to the correct number of sig figs
$3.514\ cm\ -\ 2.13\ cm\ =$	1.384 cm =	1.38 cm

Explanation: One of the quantities has only two decimal places. Therefore, the answer should be rounded to only two decimal places.

Chapter 2: Atomic Structure and the Periodic Table of Elements

- The idea of small, invisible particles being the building blocks of matter can be traced back more than 2,000 years to the Greek philosophers Democritus and Leucippus.
- These particles, considered to be so small and indestructible that they could not be divided into smaller particles, were called atoms, the Greek word for indivisible. The English word atom comes from this Greek word.
- This early concept of atoms was not based upon experimental evidence but was simply a result of thinking and reasoning on the part of the philosophers.
- It was not until the eighteenth century that experimental evidence in favor of the atomic hypothesis began to accumulate.
- Finally, around 1805, John Dalton proposed some basic assumptions about atoms based on what was known through scientific experimentation and observation at that time. These assumptions are very closely related to what scientists presently know about atoms. For this reason, Dalton is often referred to as the father of modern atomic theory. Some of these ideas:

1. All matter is made up of very small, discrete particles called atoms.
2. All atoms of an element are alike in weight, and this weight is different from that of any other kind of atom.
3. Atoms cannot be subdivided, created, or destroyed.
4. Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
5. In chemical reactions, atoms are combined, separated, or rearranged.

- By the second half of the 1800s, many scientists believed that all the major discoveries related to the elements had been made. The only thing left for young scientists was to do was to refine what was already known.
- This came to a surprising halt when J.J. Thompson discovered the electron beam in a cathode ray tube in 1897.
- Soon afterward, Henri Becquerel announced his work with radioactivity, and Marie Curie and her husband, Pierre, set about trying to isolate the source of radioactivity in their laboratory in France.
- During the late nineteenth and early twentieth centuries, more and more physicists turned their attention to the structure of the atom.
- In 1913 the Danish physicist Niels Bohr published a theory explaining the spectrum line of hydrogen. He proposed a planetary model that quantized the energy of electrons to specific orbits.
- The work of Louis de Broglie and others in the 1920s and 1930s showed that quantum theory described a more probabilistic model of where the electrons could be found that resulted in the theory of orbitals.

Basic Electric Charges

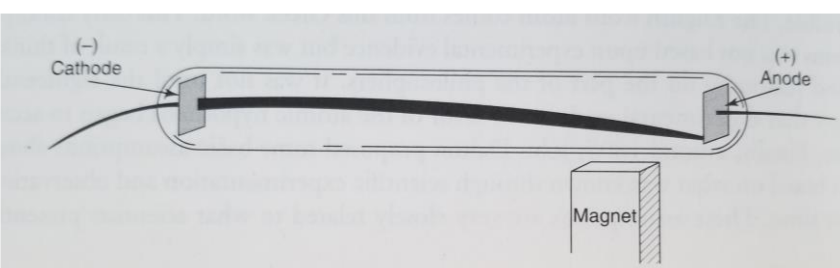


Figure 3. Cathode Ray Tube

- The discovery of the electron as the first subatomic particle is credited to J.J. Thomson (England, 1897).
- He used an evacuated tube connected to a spark coil.
- As the voltage across the tube was increased, a beam became visible. This was referred to as a cathode ray.

- Thomson found that the beam was deflected by both electric and magnetic fields.
- Therefore, he concluded that cathode rays are made up of very small, negatively charged particles, which became known as electrons.
- Further experimentation led Thomson to find the ratio of the electrical charge of the electron to its mass. This was a major step toward understanding the nature of the particle.

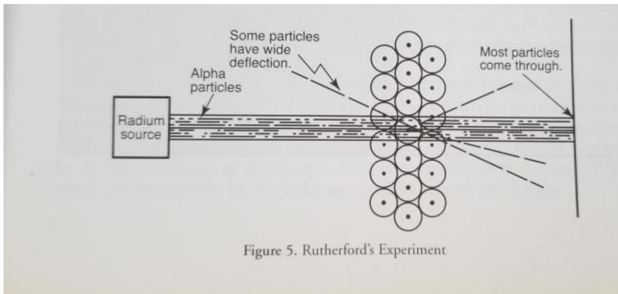


Figure 5. Rutherford's Experiment

- Ernest Rutherford (England, 1911) performed a gold foil experiment that had tremendous implications for atomic structure.
- Alpha particles passed through the foil with few deflections.
- However, some deflections were almost directly back toward the source. This was unexpected and suggested an atomic model with mostly empty space between a **nucleus**, in which most of the mass of the atom was located and which was positively charged, and the electrons that defined the volume of the atom.
- After two years of studying the results, Rutherford finally came up with an explanation. He reasoned that the rebound alpha particles must have experienced some powerful force within the atom. And he assumed this force must occupy a very small amount of space, because so few alpha particles had been deflected.
- He concluded that the force must be a densely packed bundle of matter with a positive charge. He called this positive bundle the nucleus.
- He further discovered that the volume of a nucleus was very small compared with the total volume of an atom.
- The electrons, he suggested, surrounded the positively charged nucleus like planets around the sun, even though he could not explain their motion.
- Further experiments showed that the nucleus was made up of still smaller particles called **protons**.
- Rutherford realized, however, that protons, by themselves, could not account for the entire mass of the nucleus. He predicted the existence of a nuclear particle that would be neutral and would account for the missing mass.
- In 1932, James Chadwick discovered this particle, the **neutron**.

Bohr Model of the Atom

- In 1913, Niels Bohr proposed his model of the atom. This pictured the atom as having a dense, positively charged nucleus and negatively charged electrons in specific spherical orbits, also called energy levels or **shells** around this nucleus.
- These energy levels are arranged in circles around the nucleus, and each energy level is designated by a number: 1, 2, 3, ...The closer to the nucleus, the less energy an electron needs, but it has to gain energy to go from one level to another that is father away from the nucleus.

Principal Energy Level	Max Number of Electrons ($2n^2$)

1	2
2	8
3	18
4	32
5	50

Components of Atomic Structure

Particle	Charge	Symbol	Actual Mass	Relative Mass Compared to Proton	Discovery
Electron	$-(e^-)$	$^{0}_{-1}e$	$\boldsymbol{9.109 \times 10^{-28}}$g	1/1,837	J.J. Thomson-1897
Proton	$+(p^+)$	$^{1}_{1}p$	$\boldsymbol{1.673 \times 10^{-24}}$g	1	Early 1900s
Neutron	$0(n^0)$	$^{1}_{0}n$	$\boldsymbol{1.675 \times 10^{-24}}$g	1	J.C. Chadwick-1932

- The number of protons in the nucleus of an atom determines the **atomic number**.
- All atoms of the same element have the same number of protons and therefore the same atomic number; atoms of different elements have different atomic numbers.
- Thus, the atomic number identifies the element.
- The sum of the number of protons and the number of neutrons in the nucleus is called the **mass number**.
- In some cases, different types of atoms of the same element have different masses. Protium, deuterium, and tritium are isotopes of hydrogen.
- Isotopes are atoms of the same element that have different masses. The isotopes of a particular element have the same number of protons and electrons but different number of neutrons. In all three isotopes of hydrogen, the positive charge of the single proton is balanced by the negative charge of the electron. Most elements consist of mixtures of isotopes.
- The percentage of each isotope in the naturally occurring element on Earth is nearly always the same, no matter where the element is found. The percentage at which each of an element's isotopes occurs in nature is taken into account when calculating the element's average atomic mass.
- **Average atomic mass** is the weighted average of the atomic masses of the naturally occurring isotopes of an element.

Calculating Average Atomic Mass

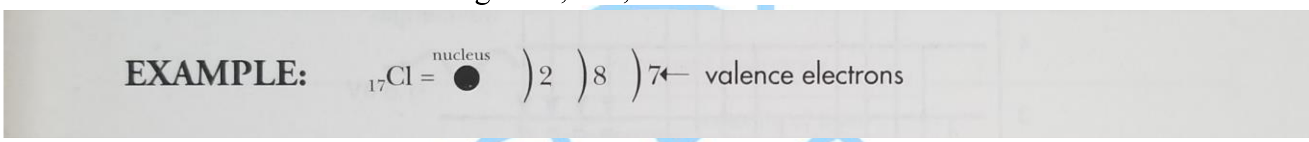
- The average atomic mass of an element depends on both the mass and the relative abundance of each of the element's isotopes.
- For example, naturally occurring copper consists of 69.17% copper-63, which has an atomic mass of 62.919 598 amu, and 30.83% copper-65, which has an atomic mass of copper can be calculated by multiplying the atomic mass of each isotope by its relative abundance (expressed in decimal form) and adding the results.

$$0.6917 \times 62.919\,598\ amu\ + 0.3083 \times 64.927\,793\ amu\ =\ 63.55\ amu$$

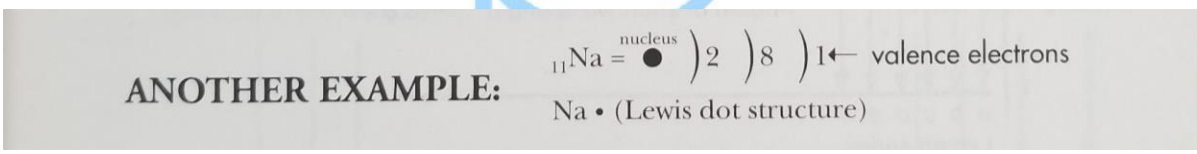
- Therefore, the calculated average atomic mass of naturally occurring copper is 63.55 amu.
- Average atomic masses of the elements listed in the Periodic Table, rounded to one decimal place for use in calculations.

Valence Electrons

- Each atom attempts to have its outer energy level complete and accomplishes this by borrowing, lending, or sharing its electrons.
- The electrons found in the outermost energy level are called **valence electrons**.
- The remainder of the electrons are called core electrons.
- The absolute number of electrons gained, lost, or borrowed is referred to as the valence of the atom.



- This picture can be simplified to :Cl:, showing only the valence electrons as dots in an electron dot notation. This is called the **Lewis structure** of the atom.
- To complete its outer orbit to eight electrons, chlorine must borrow an electron from another atom. Its valence number then is 1.
- As stated above, when electrons are gained, we assign a $-$sign to this number, so the oxidation number of chlorine is -1.



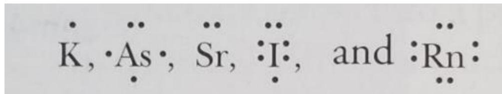
Since sodium tends to lose this electron, its oxidation number is $+1$.

Atomic Spectra

- The Bohr model was based on a simple postulate. Bohr applied to the hydrogen atom the concept that the electron can exist only in certain energy levels without an energy change but that, when the electron changes its state, it must absorb or emit the exact amount of energy that will bring it from the initial state to the final state.
- The **ground state** is the lowest energy state available to the electron.
- The **excited state** is any level higher than the ground state.

Lewis Structure (Electron Dot Notation)

- In 1916, G. N. Lewis devised the electron dot notation, which may be used in place of the electron configuration notation.
- The electron dot notation shows only the chemical symbol surrounded by dots to represent the electrons in the incomplete outer level. Examples are:



- The symbol denotes the nucleus and all electrons except the valence electrons. The dots are arranged at the four sides of the symbol and are paired when appropriate.

PERIODIC TABLE OF ELEMENTS

- The history of the development of a systematic pattern for the elements includes the work of a number of scientists.
- Dimitri I. Mendeleev in 1869 proposed a table containing 17 columns and is usually given credit for the first periodic table since he arranged elements in groups according to their atomic weights and properties.
- In 1871 Mendeleev rearranged some elements and proposed a table of eight columns, obtained by splitting each of the long periods across into a period of seven elements, an eighth group containing the three central elements, and a second period of seven elements. Mendeleev's table had the elements arranged by atomic weights with recurring properties in a periodic manner.
- The horizontal rows of the periodic table are called **periods** or **rows**. There are seven periods, each of which begins with an atom having only one valence electron and ends with a complete outer shell structure of an inert gas. The first three periods are short, consisting of 2, 8, and 8 elements, respectively. Periods 4 and 5 are longer, with 18 each, while period 6 has 32 elements, and period 7 is incomplete with 22 elements, most of which are radioactive and do not occur in nature. The vertical columns of the Periodic Table are called groups or families. The elements in a group exhibit similar or related properties. In 1984 the IUPAC agreed that the groups would be numbered 1 through 18.

PROPERTIES RELATED TO THE PERIODIC TABLE

- Metals are found on the left of the chart with the most active metal in the lower left corner.
- Nonmetals are found on the right side with the most active nonmetal in the upper right corner.
- The noble or inert gases are on the far right.
- Since the most active metals react with water to form bases, the Group 1 metals are called alkali metals.
- As you proceed to the right, the base-forming property decreases and the acid-forming properties increase.
- The metals in the first two groups are light metals, and those toward the center are heavy metals.
- The elements found along the dark line in the periodic table are called metalloids. These elements have certain characteristics of metals and other characteristics of nonmetals.

General summary statements about the Periodic Table:

- Acid-forming properties increase from left to right on the table.
- Base-forming properties are high on the left side and decrease to the right.
- The atomic radii of elements decrease from left to right across a period.
- First ionization energy increases from left to right across a period.
- Metallic properties are greatest on the left side of the table and decrease to the right.

- Nonmetallic properties are greatest on the right side of the table and decrease to the left.

Atomic Radii

1. Atomic Radii decreases from left to right across a period in the Periodic Table (until the noble gases).
    a. Since the number of electrons in the outer principal energy level increases as you go from left to right in each period, the corresponding increase in the nuclear charge b/c of the additional protons pulls the electrons more tightly around the nucleus.
2. Atomic radii increase from top to bottom in a group or family.
    a. For a group of elements, the atoms of each successive member of another outer principal energy level in the electron configuration and the electrons there are held less tightly by the nucleus. This is because of their increased distance from the nuclear positive charge and the shielding of this positive charge by all the core electrons.

Ionic Radius Compared with Atomic Radius

- Metals tend to lose electrons in forming positive ions.
- With this loss of negative charge, the positive nuclear charge pulls in the remaining electrons closer and thus reduces the ionic radius below that of the atomic radius.
- Nonmetals tend to gain electrons in forming negative ions.
- With this added negative charge, which increases the inner electron repulsion, the ionic radius is increased beyond the atomic radius.

Electronegativity

- The electronegativity of an element is a number that measures the relative strength with which the atoms of the element attract valence electrons in a chemical bond.
- This electronegativity number is based on an arbitrary scale going from 0 to 4. In general, a value of less than 2 indicates a metal.
- The electronegativity decreases down a group and increases across a period.

Ionization Energy

- Atoms hold their valence electrons, then, with different amounts of energy.
- If enough energy is supplied to one outer electron to remove it from its atom, this amount of energy is called the **first ionization energy**.
- With the first electron gone, the removal of succeeding electrons becomes more difficult because of the imbalance between the positive nuclear charge and the remaining electrons.
- The lowest ionization energy is found with the least electronegative atom.

THE NATURE OF RADIOACTIVE EMISSIONS

- Three different types of radioactive emission: alpha particles, beta particles, and gamma rays.
- Alpha particles are positively charged particles of helium nuclei, beta particles are streams of high-speed electrons, and gamma rays are high-energy radiations similar to X-rays.

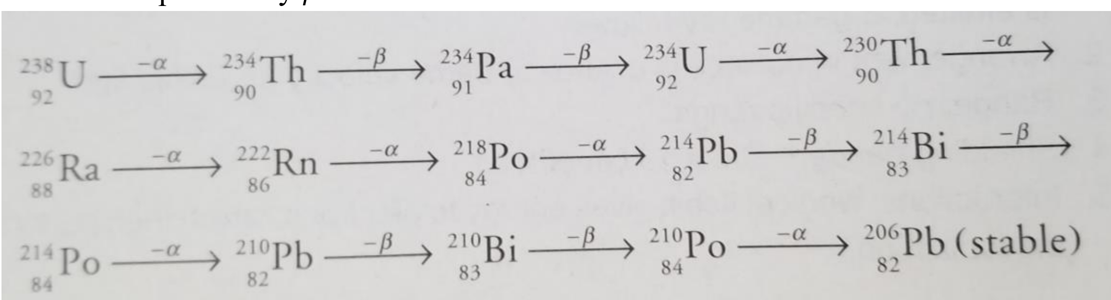
DECAY SERIES, TRANSMUTATIONS, AND HALF-LIFE

- The nuclei of uranium, radium, and other radioactive elements are continually disintegrating. It should be emphasized that spontaneous disintegration produces the gas known as radon.
- The time required for half of the atoms of a radioactive nuclide to decay is called its **half-life**.

Ex:

- Radium on average, half of the radium nuclei present will have disintegrated to radon in 1,590 years
  - In another 1,590 years half of the remainder will decay and so on
- When a radium disintegrates, it loses an alpha particle upon gaining two electrons and becomes a neutral helium atom
- The remainder becomes **radon**

- A conversion of an element to a new element (b/c of a change in the number of protons) is a **transmutation**. The following uranium-radium disintegration series shows how a radioactive atom may change when it loses each kind of article. Note that an atomic number is shown by a subscript ($_{92}U$) and the isotopic mass by a subscript (^{238}U). The alpha particle is represented by the Greek symbol, α, and the beta particle by β.



$$^{238}_{92}U \xrightarrow{-\alpha} {}^{234}_{90}Th \xrightarrow{-\beta} {}^{234}_{91}Pa \xrightarrow{-\beta} {}^{234}_{92}U \xrightarrow{-\alpha} {}^{230}_{90}Th \xrightarrow{-\alpha}$$

$$^{226}_{88}Ra \xrightarrow{-\alpha} {}^{222}_{86}Rn \xrightarrow{-\alpha} {}^{218}_{84}Po \xrightarrow{-\alpha} {}^{214}_{82}Pb \xrightarrow{-\beta} {}^{214}_{83}Bi \xrightarrow{-\beta}$$

$$^{214}_{84}Po \xrightarrow{-\alpha} {}^{210}_{82}Pb \xrightarrow{-\beta} {}^{210}_{83}Bi \xrightarrow{-\beta} {}^{210}_{84}Po \xrightarrow{-\alpha} {}^{206}_{82}Pb \text{ (stable)}$$

Radioactive Decay and Nuclear Charge

Type of Decay	Decay Particle	Particle Mass	Particle Charge	Change in Nucleon Number	Change in Atomic Number
Alpha Decay	α	4	2+	Decreases by 4	Decreases by2
Beta Decay	β	0	1-	No Change	Increases by 1
Gamma Radiation	γ	0	0	No Change	No Change
Positron emission	β^+	0	1+	No Change	Decreases by 1
Electron capture	e^-	0	1-	No Change	Decreases by 1

Nuclear Systems for Subatomic Particles

Particle	Symbols
Proton	p

Neutron	n
Electron	e^- or β^-
Positron	e^+ or β^+
Alpha Particle	α
Beta Particle	βor β^-
Gamma Ray	γ

RADIOACTIVE DATING

- A helpful application of radioactive decay is in the determination of the ages of substances such as rocks and relics that have bits of organic material trapped in them.
- Because carbon-14 has a half-life of about 5,700 years and occurs in the remains of organic materials, it has been useful in dating these materials. A small percentage of CO_2in the atmosphere contains carbon-14. The stable isotope of carbon is carbon-12. Carbon-14 is a beta emitter and decays to form nitrogen-14:

- In any living organism, the ratio of carbon-14 to carbon-12 is the same as in the atmosphere b/c of the constant interchange between organisms and surroundings.
- When an organism dies, this interaction stops, and the carbon-14 gradually decays to nitrogen. By comparing the relative amounts of carbon-14 and carbon-12 in the remains, the age of the organism can be established.

NUCLEAR REACTIONS

Nuclear fission reactions are the splitting of a heavy nucleus into two or more lighter nuclei.

- A nuclear chain reaction is a reaction in which an initial step, such as the reaction above, leads to a succession of repeating steps that continues indefinitely.
- A nuclear fission reaction is the combination of very light nuclei to make a heavier nucleus.
- Extremely high temperatures and pressures are required in order to overcome the repulsive forces of the two nuclei.

- The energy released in a nuclear reaction (either fission or fusion) comes from the fractional amount of mass converted into energy.
- Nuclear changes convert matter into energy. Energy released during nuclear reactions is much greater than the energy released during chemical reactions.

Chapter 3: BONDING

- A **molecule** is defined as the smallest particle of an element or a compound that retains the characteristics of the original substance.
- When atoms do combine to form molecules, there is a shifting of valence electrons, the electrons in the outer energy level of each atom. This more stable form may be achieved by the gain or loss of electrons or the sharing pairs of electrons.
- The resulting attraction of the atoms involved is called a **chemical bond**. When a chemical bond forms, energy is released; when this bond is broken, energy is absorbed.

TYPES OF BONDS

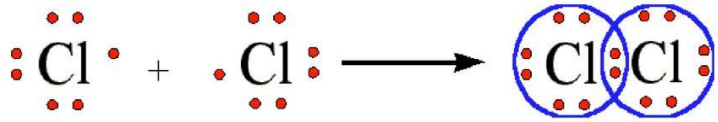
Ionic Bonds

- When the electronegativity values of two kinds of atoms differ by 1.7 or more (especially differences greater than 1.7), the more electronegative atom will borrow the electrons it needs to fill its energy level, and the either atom will lend electrons until it has a complete energy level.
- Because of this the borrower becomes negatively charged and is called an anion; the lender becomes positively charged and is called a cation. They are now referred to as **ions**, and the bond or attraction b/w them is called an **ionic bond**.

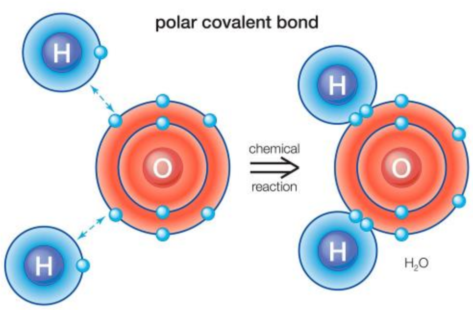
- These ions do not form an individual molecule in the liquid or solid phase but are arranged into a crystal lattice containing many such ions.
- Ionic solids of this type tend to have high melting points and will not conduct a current of electricity until they are in the molten state.

Covalent Bonds

Nonpolar Covalent Bond: When the electronegativity difference b/w 2 or more atoms is 0 or very small (not greater than 0.4), the atoms tend to share the valence electrons in their respective outer energy levels.



- These covalent bonded molecules do not have electrostatic charges like those of ionic bonded substances.
- In general, covalent compounds are gases, liquids having fairly low boiling points, or solids that melt at relatively low temperatures.
- Unlike ionic compounds, they do not conduct electric currents.



- When the electronegativity difference is b/w 0.4 and 1.6, there will not be an equal sharing of electrons b/w atoms involved. The shared electrons will be more strongly attracted to the atom of greater electronegativity.
- As the difference in the electronegativities of the two elements increases above 0.4, the polarity or degree of ionic character increases. At a difference of 1.7 or more, the bond has more than 50% ionic character.

Polar Covalent Bond: When the difference is between 0.4 and 1.6

- When these non symmetrical polar bonds are placed around a central atom, the overall molecule is polar. Both the bonds and the molecules could be described as **polar**.
- Polar molecules are also often referred to as dipoles b/c the whole molecule itself has 2 distinct ends from a charge perspective. However, polar covalent bonds exist in some nonpolar molecules. Examples are CO_2, CH_4, CCl_4.
- In the **covalent** bonds described so far, the shared electrons in the pair contributed one each from the atoms bonded. In some cases, however, both electrons from the shared pair are supplied by only one of the atoms. Two examples are the bonds in NH_4+ and H_2SO_4.

Metallic Bonds

- In most metals, one or more of the valence electrons become detached from the atom and migrate in a "sea" of free electrons among the positive metal ions.
- The attractive force strength varies with the nuclear positive charge of the metal atoms and the number of electrons in this electron sea.
- Both of these factors are reflected in the amount of heat required to vaporize the metal.
- The strong attraction b/w these differently charged particles forms a **metallic bond**. Because of this firm bonding, metals usually have high melting points, show great strength, and are good conductors of electricity.

INTERMOLECULAR FORCES OF ATTRACTION

The term **intermolecular forces** refers to attractions b/w molecules. It is proper to refer to all intermolecular forces as **van der Waals forces**.

Dipole-Dipole Attraction

- One type of van der Waals forces is dipole-dipole attraction.
- The asymmetrical distribution of electric charges leads to positive and negative charges in the molecules, referred to as **dipoles**.
- In polar molecular substances, the dipoles line up so that the positive pole of one molecule attracts the negative pole of another.
- The force of attraction b/w polar molecules is called dipole-dipole attractions.

London Dispersion Forces

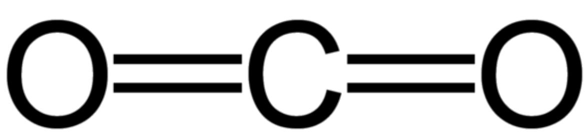
- Another type of van der Waals force is allied London dispersion forces.
- Found in both polar and nonpolar molecules, a molecule/atom that usually is nonpolar sometimes becomes polar b/c the constant motion of its electrons may cause uneven charge distribution at any one instant.
- When this occurs, the molecule/atom has a temporary dipole. This dipole can then cause a second, adjacent atom to be distorted and to have its nucleus attracted to the negative end of the first atom.
- London dispersion forces are the weakest of all the electrical forces b/w atoms or molecules.
- They help to explain why nonpolar substances (nonpolar substances, halogens) condense into liquids and then freeze into solids when the temperature is lowered sufficiently.
- They also help to explain why liquids composed of discrete molecules w/ no permanent dipole attraction have low boiling points relative to their molecular masses.
- Compounds in the solid state that are bound mainly by this type of attraction have rather soft crystals, are easily deformed, and vaporize easily.
- B/c of the low IMF, the melting points are low and evaporation takes place so easily that it may occur at room temperature.

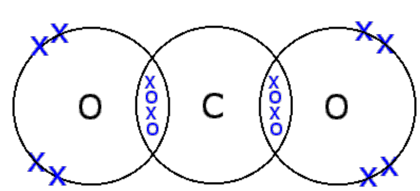
Hydrogen Bonds

- A hydrogen bond is when a hydrogen atom is bonded to a highly electronegative atom, itss positive charge will have an attraction for neighboring electron pairs.
    - A proton or hydrogen nucleus has a high concentration of positive charge.
- The more strongly polar the molecule is, the more effective the hydrogen bonding is in binding the molecules into a larger unit.
    - As a result, the boiling points of such molecules are higher than those of similar nonpolar molecules.
- It explains why some substances have unexpectedly low vapor pressures, high heats of vaporization, and high melting points. In order for vaporization or melting to take place, molecules must be separated.
- Energy must be expended to break hydrogen bonds and thus break down large clusters of molecules into separate molecules.

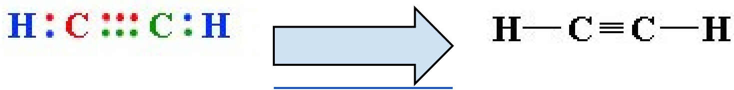
DOUBLE AND TRIPLE BONDS

- To achieve the octet structure, an outer energy level resembling the noble gas configuration of 8 electrons, it is necessary for some atoms to share 2 or even 3 pairs of electrons.
- Sharing 2 pairs of electrons produces a double bond.





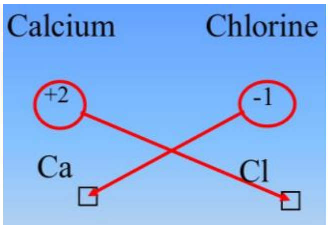
- In the line formula, only the shared pair of electrons is indicated by a bond (–).
- The sharing of 3 electron pairs results in a triple bond.



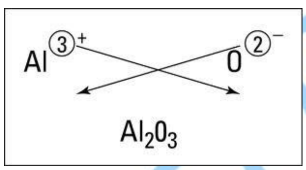
Chapter 4: Chemical Formulas

NAMING AND WRITING CHEMICAL FORMULAS

- Category 1: Binary ionic compounds where the metal present forms only a single type of positively charged ion (cation)
    - Metallic ions from groups 1 and 2 of the Periodic Table
    - These metallic ions have only one type of charge
    - Composed of a positive ion (cation) that is written first and then a negative ion (anion0
        - Ex: $CaCl_2$
            1. Name the cation first and then the anion.
            2. The one-atom cation takes its name from the element. Therefore, the calcium ion, Ca^{2+}, is called calcium and its chemical symbol appears first.
            3. The one-atom anion with which the cation combines is named by taking the root of the element's name and adding -ide. The anion's name comes second. Therefore, thee chlorine ion, Cl^{-}, is called chloride.
            4. The name of this compound is calcium chloride.
- A quick way to determine the formula of a binary ionic compound is to use the **criss-cross rule**.



- Example 1: To determine the formula for calcium chloride, first write the ionic forms with their associated charges.
- Next, move the numerical value of the metal ion's superscript (without the charge) to the subscript of the nonmetal's symbol. Then take the numerical value of the nonmetal's superscript and make it the subscript of the metal as shown above.
- NOTE: The numerical value 1 is not shown in the final formula.
- Now the chlorine's 1 as the subscript of the calcium and the calcium's 2 as the subscript of the chloride. As a result you have $CaCl_2$as the final formula for calcium chloride.
- Example 2: Write the name and formula for the product formed when aluminum reacts with oxygen. First write the name.
    1. Name the cation first and then the anion.
    2. The one-atom cation takes its name from the name of the element. Therefore, the aluminium ion, Al^{3+}, is called aluminum and its chemical symbol appears first.
    3. The one-atom anion w/ which the cation combines is named by taking the root of the element's name and adding -ide. The anion's name comes second. Therefore, the oxygen ion, O^{2-}, is called oxide.
    4. The name of this compound is aluminum oxide.



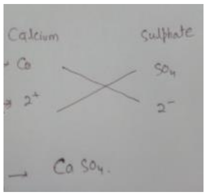
To determine the formula for aluminum oxide, first write the ionic forms with their associated charges.
Next, move the numerical value of the Al's superscript (w/o the charge) to the subscript of the O symbol. Do the same w/ the 2 of the O. Now the 2 is the subscript of the aluminum and the 3 as a subscript of the oxygen. You now have Al_2O_3.

- Category 2: Binary ionic compounds where the metal forms more than one type of ionic compounds with a given negatively charged ion (anion)
    - The metals form more than one ion, each w/ a diff charge
    - The metallic ions (cation) ionically bind w/ a negatively charged ion (anion)
        - Example: The compound containing the Fe^{2+}ion and the compound containing the Fe^{3+}ion both combine w/ chloride ion to form 2 different compounds. You get the formula $FeCl_2$ for iron(II) chloride.
- The compound formed using the Fe^{3+}ion and the chloride ion is $FeCl_3$, which is iron (III) chloride.
- The names iron (II) chloride and iron (III) chloride are arrived at by using the Roman numerals in parentheses to indicate the charge of the metallic ion used as the cation.

Category 1 and 2 Ionic Compounds Formed W/ Polyatomic Ions
- A polyatomic ion is a group of elements that act like a single ion when forming a compound.
- The bonds within are predominantly covalent. However, the group as a whole has an excessive charge, usually negative, b/c of an excess of electrons.
- If the compounds formed w/ the polyatomic ions consist of 3 elements, they are called ternary compounds.
- Sometimes an element combines w/ oxygen to form more than just 2 polyatomic ions, the prefix hypo- is used to name the polyatomic ion w/ the fewest oxygen ions and the prefix per- to name the polyatomic ion w/ the most oxygen ions.

Writing Formulas w/ Polyatomic Ions
- Simply treat the polyatomic ion as if it were a single anion.

- If the cation is from category 1, follow the rules for category 1. If the cation is from category 2, follow the rules for category 2.



- Example 1: Use the crisscross method to write the formula for calcium sulfate, a category 1 cation and a polyatomic ion.
- The final formula is $CaSO_4$. Notice the subscripts "2" are omitted.
- Category 3: Binary covalent compounds formed b/w 2 nonmetals
  - Binary covalent compounds are formed b/w 2 nonmetals
    1. The first name element in the formula is named first, using the full elemental name.
    2. The second element is named as if it were an anion and uses its elemental name.
    3. Prefixes are used to denote the number of the second element present. These prefixes are shown in the table below.
    4. The prefix mono- is never used for naming the first element. For example, CO is called carbon monoxide, not monocarbon monoxide.

Prefix	Number	Prefix	Number
mono-	1	hexa-	6
di-	2	hepta-	7
tri-	3	octa-	8
tetra-	4	nona-	9
penta-	5	deca-	10

- To write the formula for binary covalent compounds, use the same steps as when writing the formula of ionic compounds.
  1. The symbol of the first element in the formula is written first, followed by the second element.
  2. Use the prefix(es) denoted in the name for the number of each element present in the formula.
     Examples: Sulfur hexafluoride (SF_6) and Phosphorus trichloride (PCl_3)

OXIDATION STATES AND FORMULA WRITING
- An oxidation state is assigned to each member of a formula or polyatomic ion
- It's designated by a small, whole-number superscript preceded by a plus or minus sign
- This is not to be confused with the ionic charges used to the right of ionic charge
- These charges are directly related to the bonding that occurs in compounds
- They're also used to track electron transfers

The Rules for Assigning an Oxidation State
- Remember: The sum of the oxidation states must be zero for an electrically neutral compound
- For an ion, the sum of the oxidation states must equal the charge on the ion
1. The oxidation state of an atom in an element is zero.
2. The oxidation state of a monatomic ion is the same as its charge.
3. The oxidation state of fluorine is -1 in compounds.

4. The oxidation state of oxygen is usually -2 in its compounds.
5. The oxidation state of hydrogen in most compounds is +1.

EXAMPLE 1: In Na_2SO_4, what is the oxidation state of sulfur?

Na_2SO_4

We know the oxidation state of one atom of Na is +1. There are two atoms of Na. Using rule 2, $2 \times (+1) = +2$

We know the oxidation state of one atom of O is −2. There are four atoms of O. Using rule 4, $4 \times (-2) = -8$

Since the positive sum plus the negative sum must equal 0,

$$(+2) + x + (-8) = 0$$

The sulfur must have a +6 oxidation state.

EXAMPLE 2: What is the oxidation state of chromium in $K_2Cr_2O_7$?

$$K = 2 \times (+1) = +2$$
$$Cr = 2 \times (x) = 2x$$
$$O = 7 \times (-2) = -14$$
$$(+2) + (2x) + (-14) = 0$$
$$2x = +12$$
$$x = +6$$

The chromium has a +6 oxidation state.

In a polyatomic ion, the algebraic sum of the positive and negative oxidation states of all the atoms in the formula must be equal to the charge of the ion.

CHEMICAL FORMULAS: THEIR MEANING AND USE

- By using the atomic masses assigned to the elements, we can find the **formula mass** of a compound.
- The formula mass is determined by multiplying the atomic mass units (as a whole number) by the subscript for that element and then adding these values for all the elements in the formula.
- If the formula represents the actual makeup of one molecule of the substance, **molecular mass** can be used as well.
- The simplest ratio formula is called the empirical formula, and the actual formula is called the true formula.

Example 1: $Ca(OH)_2$(one calcium amu + two hydrogen amu = formula mass)

1Ca (amu = 40) = 40
2O (amu = 16) = 32
2H (amu = 1) = 2.0

Formula Mass $Ca(OH)_2$= 74 amu

- If you have 6.02×10^{23} atoms of an element, then the atomic mass units can be expressed in grams, and then the formula mass can be called the molar mass.

Example 2: Fe_2O_3

2Fe (amu = 56) = 112

3O (amu = 16) = 48.0

Formula mass Fe_2O_3= 160. Amu

- Finding the **percentage composition** is finding what percent of the total weight of a compound is made up of a particular element.

$\frac{Total\ amu\ of\ the\ element\ in\ the\ compound}{Total\ formula\ amu} \times 100\%$ =Percentage composition of that element

To find the percent composition of calcium in calcium hydroxide:

$\frac{Ca = 32\ amu}{Formula\ mass = 74\ amu} \times 100\% = 54\%$Calcium

Example: Find the percent compositions of Cu and H_2Oin the compound $CuSO_4 \cdot 5H_2O$.

First, calculate the formula mass:

1 Cu = 64 amu

1 S = 32 amu

4 O = 64(4×16) amu

$5H_2O$= 90. (5×18) amu

---

250 amu

And then find the percentages:

Percentage Cu: $\frac{Cu = 64\ amu}{Formula\ mass = 250\ amu} \times 100\% = 26\%$ Percentage H_2O:$\frac{5.0\ H_2O = 90.amu}{Formula\ mass = 250\ amu} \times 100\% = 36\%$

When you are given the percentage of each element in a compound, you can find the empirical formula as shown w/ the following example:

Given Ba = 58.81%, S = 13.73%, and O = 27.46%. Find the empirical formula.

1. Divide each percent by the amu of the element.
   a. Ba: 58.8/137 = 0.43, S:13.7/32 = 0.43, O: 27.5/16 =1.72
2. Manipulate numbers to get small whole numbers. Try dividing them all by the smallest first. In this case, divide each by 0.43.
   a. Ba: 0.43/0.43 = 1, S: 0.43/0.43 = 1, O: 1.72/0.43 = 4
3. The formula is $BaSO_4$.

WRITING AND BALANCING EQUATIONS

- An equation is a simplified way of recording a chemical change.
- Chemical symbols and formulas are used to represent the reactants and the products.

Example: $H_2 + O_2 \rightarrow H_2O$

- While the left side has 2 atoms of oxygen, the right side only has one. To get the number of atoms of each element represented on the left side to equal the number on the right.
- To do this, we can only use coefficients in front of the formula.
- THE SUBSCRIPTS IN THE FORMULA MAY NOT BE CHANGED
- If we put 2 in front of the $H_2O$the number of oxygen atoms represented on the 2 sides of the equation are equal. However, there are now 4 hydrogens on the right side w/ only 2 on the left.This can be corrected by using a coefficient of 2 in front of the H_2 .

$$2H_2 + O_2 \rightarrow 2H_2O$$

SHOWING PHASES IN CHEMICAL EQUATIONS

- Once an equation is balanced, you can indicate the phases of substances, telling whether each substance is in the liquid phase (l), the gaseous phase (g), or the solid phase (s).
- Since many solids will not react to any appreciable extent unless they are dissolved in water, the notation (aq) is used to indicate the substance exists in a water (aqueous) solution.

Ex: (aq) + Zn (s) → $ZnCl_2$ (aq) + H_2(g)

WRITING IONIC EQUATIONS

- Ionic equations are used to show only the substances that react in the chemical action and b/c they stress the reaction and production of ions

Reactants:

2NCl(aq) releases → $2H^+$(aq) + $2Cl^-$(aq) in solution
Zn(s) stay as → Zn(s) particles

Products:

$ZnCl_2$(aq) releases as → Zn^{2+}(aq) + $2Cl^-$(aq) in solution
H_2(g) stay as → H_2(g) molecules

The complete equation: $2H^+(aq) + 2Cl^-(aq) + Zn\ (s) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq) + H_2(g)$

Chapter 5: Gases and the Gas Laws

GAS LAWS AND RELATED PROBLEMS

Charles's Law ($\frac{V}{T} = K$)

- Jacques Charles discovered that when a gas under constant pressure is heated from $0^\circ C$ to $1^\circ C$, it expands 1/273 of its volume.
- If a gas at $0^\circ C$ was cooled to $-273^\circ C$, its volume would be zero.
- If the pressure remains constant, the volume of a gas varies directly as the absolute temperature. Then:
  - Initial $\frac{V_1}{T_1}$ =Final $\frac{V_2}{T_2}$ at constant pressure or $\frac{V}{T} = K$

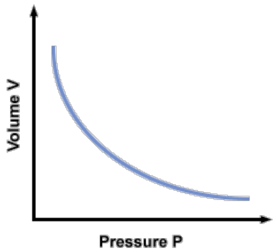
Example:The volume of a gas at $20^\circ C$ is 500 mL. Find its volume at standard temperature if pressure is held constant.

Constant Temperatures:

$20^\circ C = 20^\circ + 273 = 293\ K$
$0^\circ C = 0^\circ + 273 = 273\ K$

Use the formula and substitute known values.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
$\frac{500.mL}{293} = \frac{x\ mL}{273}$
$x\ mL = 465\ mL$



Boyle's Law (PV = K)

- Robert Boyle, found that the volume of a gas decreases when the pressure on it is increased, and vice versa, when the temperature is held constant.
- If the temperature remains constant, the volume of a gas varies inversely as the pressure changes. Then:
- $P_1V_1 = P_2V_2$ at a constant temperature or PV = K

Ex: Given the volume of a gas as 200. mL at 1.05 atm pressure, calculate the volume of the same gas at 1.01 atm. Temperature Is held constant.

- $P_1V_1 = P_2V_2$
- $V_2 = V_1 \times \frac{P_1}{P_2}$
- $= 200.mL \times \frac{1.05\ atm}{1.01\ atm} = 208\ mL$

Combined Gas Law

- A combination of the 2 previous gas laws:
  - $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Ex: The volume of a gas at 780. mm pressure and $30^\circ C$ is 200. mL. What volume would the gas occupy at STP?

- $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- $V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1}$
- $V_2 = 500.mL \times \frac{780.mm\ Hg}{760.mm\ Hg} \times \frac{273\ K}{303\ K} = 462\ mL$

Gay-Lussac's Law

- At constant volume, the pressure of a given mass of gas varies directly w/ the absolute temperature. Then:
  - $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ at constant temperature or $\frac{P_1}{T_1} = k$

Ex: A steel tank contains a gas at $27^\circ C$ and a pressure of 12.0 atms. Determine the gas pressure when the tank is heated to $100^\circ C$.

- $\frac{P_1}{T_1} = \frac{P_2}{T_2}$
- $P_2 = P_1 \times \frac{T_2}{T_1}$
- $P_2 = 12.0\ atm \times \frac{373\ K}{300.K} = 14.9\ atm$

Chapter 6: Stoichiometry (Chemical Calculations) and the Mole Concept

THE MOLE CONCEPT

- A mole describes a quantity of particles, 1 mole = 6.02×10^{23} particles
- The number 6.02×10^{23} is also often referred to as Avogado's number

MOLAR MASS AND MOLES

- When the formula mass of an ionic compound is determined by the addition of its component relative atomic masses and expressed in grams, it is called the molar mass.

Ex: Find the molar mass of $CaCO_3$.

1 Ca = 40
1 C = 12
3 O = 48(3×16)

$CaCO_3 = 100.$formula mass

100. g/mol is the molar mass

- The molar mass is equivalent to the mass of 1 mole of that compound expressed in grams.

MOLE RELATIONSHIPS

- When dealing w/ elements-
  - Moles of an element ×molar mass (atomic) = mass of the element
  - Mass of an element/molar mass (atomic) = moles of the compound
- When dealing w/ compounds-
  - Moles of a compound ×molar mass (molecular) = mass of the compound
  - Mass of a compound/molar mass (molecular) = moles of the compound
- When dealing w/ the molecules of a compound
  - Moles of molecules × 6.02×10^{23}= number of molecules
  - Number of molecules/6.02×10^{23}= moles of molecules
- When dealing w/ the atoms of elements-
  - Moles of atoms × 6.02×10^{23}= number of atoms
  - Number of atoms/6.02×10^{23}= moles of atoms

DENSITY AND MOLAR MASS

- Since the density of a gas is usually given in grams/liter of gas at STP, we can use the molar volume to molar mass relationship to solve the following types of problems.

Ex: Find the molar mass of a gas when the density is given as 1.25 grams/liter.

- B/c we know that 1 mole of a gas occupies 22.4 liters at STP, we can solve this problem by multiplying the mass of 1 liter by 22.4 liters/mole.

$$\frac{1.25\ g}{L} \times \frac{22.4\ L}{1\ mol} = 28.0\ g/mol$$

MASS -VOLUME RELATIONSHIPS

Ex: How many liters of oxygen (STP) can you prepare from the decomposition of 42.6 grams of sodium chlorate?

Step 1: Write the balanced equation for the reaction.

$2\ NaClO_3 \rightarrow 2\ NaCl\ + 3O_2$

Step 2: Write the given quantity and the unknown quantity above the appropriate substances.

42.6 g x L

$2\ NaClO_3 \rightarrow 2\ NaCl\ + 3O_2$

Step 3: Calculate the equation mass or volume of each substance that has something indicated above it, and write the results under the substances.

42.6 g x L

$2\ NaClO_3 \rightarrow 2\ NaCl\ + 3O_2$

213 g 67.2 L

(2×molar (3×22.4 L)

Mass of $NaClO_3$)

StEP 4: Form the proportion.

$\frac{42.6\ g}{213\ g} = \frac{x\ L}{67.2\ L}$

Step 5: Solve for x.

x = 13.4 L O_2

- Another method to proceed from Step 3 is called dimensional analysis or the factor-label method

Step 4: $42.6\ g\ NaClO_3 \times \frac{67.2\ L\ O_2}{213\ g\ NaClO_3} = 13.4\ L\ O_2$

- Another method of solving this problem is called the mole method.

Step 3: Determine how many moles of substance are given.

$42.6 \div \frac{106.5\ g}{1\ mol\ NaClO_3} = 0.400\ mol\ NaClO_3$

The equation shows that 2 moles of $NaClO_3$ makes 3 moles O_2. Then 0.4 mole $NaClO_3$ will yield

$0.400\ mol\ NaClO_3 \times \frac{3\ mol\ O_2}{2\ mol\ NaClO_3} = 0.600\ mol\ O_2$

Step 4: Convert the moles of O_2 to liters.

$0.600\ mol\ O_2 \times \frac{22.4\ L\ O_2}{1\ mol\ O_2} = 13.4\ L\ O_2$

MASS-MASS PROBLEMS

Ex: What mass of oxygen can be obtained from heating 100. grams of potassium chlorate?

Step 1: Write the balanced equation for the reaction.

$2KClO_3 \rightarrow 2KCl + 3O_2$

Step 2: Write the given quantity and the unknown quantity above the appropriate substances.

100. g x g

$2KClO_3 \rightarrow 2KCl + 3O_2$

Step 3: Calculate the equation mass for each substance that has something indicated above it, and write the results under the substances.

100. g x g

$2KClO_3 \rightarrow 2KCl + 3O_2$

245 g 96.0 g

Using the proportion method:

Step 4: Form the proportion.

$\frac{100.g}{245\ g} = \frac{x\ g}{96.0\ g}$

Step 5: Solve for x.

x = 39.3 g of O_2

Using the factor-label method:

Step 4: The equation indicates that 245 g $KClO_3$ yields 96 g of O_2. Therefore multiplying the given quantity by a factor made up of these two quantities arranged appropriately so that the units of the answer remain uncanceled will give the answer to the problem.

$100.\ g\ KClO_3 \times \frac{96\ g\ O_2}{245\ g\ KClO_3} = 39.3\ g\ O_2$

Using the mole method:

Step 3: Determine how many moles of substances are given.

$100.\ g\ KClO_3 \times \frac{1\ mol\ KClO_3}{122.5\ g\ KClO_3} = 0.815\ mol\ KClO_3$

The equation shows that 2 moles of $KClO_3$ yields 3 moles of O_2, so 0.815 mole of $KClO_3$ will yield

$0.815\ mol\ KClO_3 \times \frac{3\ mol\ O_2}{2\ mol\ KClO_3} = 1.22\ mol\ O_2$

Step 4: Convert the moles of O_2 to grams of O_2.

$1.22\ mol\ O_2 \times \frac{32.0\ g\ O_2}{1\ mol\ O_2} = 39\ g\ O_2$

VOLUME-VOLUME PROBLEMS

Ex: What volume of NH_3 is produced when 22.4 liters of nitrogen are made to combine w/ a sufficient quantity of hydrogen under the appropriate conditions?

Step 1: Write the balanced equation for the reaction.

$N_2 + 3H_2 \rightarrow 2NH_3$

Step 2: Write the given quantity and the unknown quantity above the respective substances.

22.4 L x L

$N_2 + 3H_2 \rightarrow 2NH_3$

Using the proportion method:

Step 3: Set up a proportion using the coefficient of the substances that have something indicated above them as denominators.

$\frac{22.4\ L}{1\ L} = \frac{x\ L}{2\ L}$

Step 4: Solve for x.

x = 44.8 L

Using the factor-label method:

Step 3: The equation shows that 1 volume of N_2will yield 2 volumes of NH_3. Therefore multiplying the given quantity by a factor made up of these 2 quantities appropriately arranged so that the units of the answer remain uncanceled will solve the problem.

$22.4\ L\ N_2 \times \frac{2\ L\ NH_3}{1\ L\ N_2} = 44.8\ L\ NH_3$

Using the mole method:

Step 3: The given quantity is converted to moles.

$22.4\ L\ N_2 \times \frac{1\ mol\ N_2}{22.4\ L\ N_2} = 1\ mol\ N_2$

The equation shows that 1 mole of N_2will yield 2 moles of NH_3. Therefore using this relationship will yield

$1\ mol\ N_2 \times \frac{2\ mol\ NH_3}{1\ mol\ N_2} = 2\ mol\ NH_3$

Convert the moles of $NH_3$to liters of NH_3.

$2\ mol\ NH_3 \times \frac{22.4\ L\ NH_3}{1\ mol\ NH_3} = 44.8\ L\ NH_3$

Chapter 7: Liquids, Solids, and Phase Changes

LIQUIDS

- A form of matter that has a definite volume and takes the shape of its container.
- A liquid expands and contracts only very slightly w/ a change in temperature and lacks the compressibility typical of gases.

Surface Tension

- Molecules at the surface of a liquid experience attractive forces downward, toward the inside of the liquid, and sideways, along the surface of the liquid. On the other hand, molecules in the center of the liquid experience uniformly distributed attractive forces. This imbalance of forces at the surface of a liquid results in a property called **surface tension**.

PHASE EQUILIBRIUM

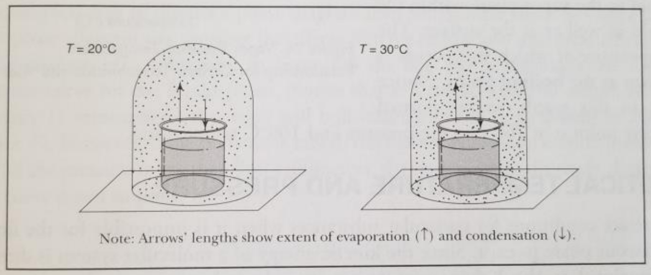


Figure 23. Closed System in Dynamic Equilibrium

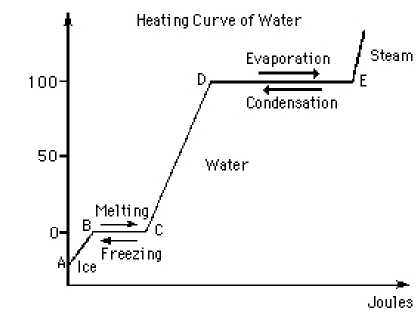
- Observation of this closed system would show an initial small drop in the water level, but after some time the level would be constant.
- At first, more energetic molecules near the surface are escaping into the gaseous phase faster than some of the gaseous water molecules are returning to the surface and possibly being caught by the attractive forces that will retain them in the liquid phase.
- After some time the rates of evaporation and condensation equalize, known as **phase equilibrium**.
- In a closed system like this, when opposing changes are taking place at equal rates, the system is said to have dynamic equilibrium.
- At higher temperatures, since the number of molecules at higher energies increases, the number of molecules in the liquid phase will be reduced and the number of molecules in the gaseous phase will be increased. The rates of evaporation and condensation will again become equal.
- **Le Chatelier's Principle**: When a system at equilibrium s disturbed by the application of a stress (a change in temperature, pressure, or concentration), it reacts so as to minimize the stress and attain a new equilibrium position.
- Heat + $H_2O(\text{l}) \leftrightarrows H_2O(\text{g})$
    - The equation shifts to the right (any similar system that is endothermic shifts to the right when temperature is increased) until equilibrium is reestablished at the new temperature.
- The molecules in the vapor that are in equilibrium w/ the liquid at a given temperature exert a constant pressure, known as equilibrium vapor pressure.

BOILING POINT
- When a liquid is heated in an open container, the liquid and vapor are not in equilibrium and the vapor pressure increases until it becomes equal to the pressure above the liquid.
- At this point the average kinetic energy of the molecules is such that they are rapidly converted from the liquid to the vapor phase within the liquid as well as the surface, the temperature at which this occurs is known as the **boiling point**.

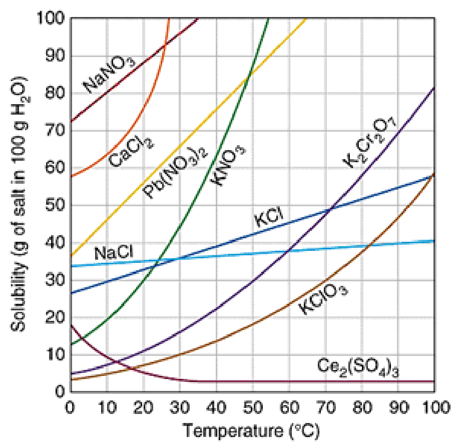
SOLIDS
- Particles are fixed in rather definite positions and maintain definite shapes.
- When heated at certain pressures, some solids vaporize directly w/o passing through the liquid phase, called **sublimation**.
- The temperature at which atomic or molecular vibrations of a solid become so great that the particles break free from fixed positions and begin to slide freely over each other in a liquid state called the **melting point**.
- The amount of energy required at the melting point temperature to cause the change of phase to occur is called the heat of fusion.

Water Calorimetry Problems

- A calorimeter is a container well insulated from outside sources of heat or cold so that most of its heat is contained in the vessel.
- From this graph, heat is being used at $0^{\circ}C$ and $100^{\circ}C$ to change the state of water, but not its temperature.
- One gram of ice at $0^{\circ}C$ needs 80 calories or 3.34×10^{2} joules to change the water at $0^{\circ}C$, called **heat of fusion**.
- Energy is being used at $100^{\circ}C$ to change water to steam, not to change the temperature.
- One gram of water at $100^{\circ}C$ absorbs 540 calories or 2.26×10^{3} joules of heat to change 1 gram of steam at $100^{\circ}C$, called **heat of vaporization**.



SOLUBILITY



- Water is referred to as "the universal solvent" b/c of the number of common substances that dissolve in water.
- When substances are dissolved in water to the extent that no more will dissolve at that temperature, the solution is said to be **saturated**.
- The substance dissolved is called the **solute** and the dissolving medium is called a **solvent**.
- As a soluble solute is added to water at a given temperature, the solute will continue to go into solution until the water cannot dissolve anymore (sloped portion of graph).
- From that point on, the solute will fall to the bottom and does not dissolve, b/c an equilibrium has established b/w molecules leaving and entering the solid phase (line on the graph becomes horizontal).
- This solution is holding the maximum amount of dissolved solute in this amount of water and at $20^{\circ}C$, this is a **saturated solution**.
- A solution that contains less solute than a saturated solution under the existing conditions is described as an **unsaturated solution**.

Chapter 8: Chemical Reactions and Thermochemistry

- One type of reaction is **combination**, or it can also be called **synthesis**. This means the formation of a compound from the union of its elements.
  - $Zn(s) + S(s) \rightarrow ZnS(s)$
  - $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$
  - $C(s) + O_2(g) \rightarrow CO_2(g)$
- The second type of reaction, **decomposition**, can also be referred to as **analysis**. This means the breakdown of a compound to release its components as individual elements or other compounds.
  - $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
  - $C_{12}H_{22}O_{11}(s) \rightarrow 12C(s) + 11H_2O(l)$
  - $2HgO(s) \rightarrow 2Hg(s) + O_2(g)$
- The 3rd type of reaction is called **single replacement** or **single displacement**. This can best be shown in which one substance is displacing another.
  - $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$
  - $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
  - $Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$
- The last type is calle **double replacement** or **double displacement** b/c there is an actual exchange of "partners" to form new compounds.
  - $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
  - $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

PREDICTING REACTIONS

1. Combination
   - A heat of formation table gives the number of kj evolved or absorbed when a mole of the compound in question is formed by the direct union of its elements.
     - A positive number indicates that heat is absorbed, a negative number when heat is evolved.
     - If the heat of formation is a large number preceded by a minus sign, the combination is likely to occur spontaneously and the reaction is exothermic.
     - If the number is small and negative or is positive, heat will be needed to get the reaction to proceed.
     - The symbol $\Delta H$is used to indicate heat of formation.

Ex: $Zn(s) + S(s) \rightarrow ZnS(s) + 202.7\ kJ \qquad \Delta H = -202.7\ kJ$

This means that 1 mole of zinc (65 grams) reacts w/ 1 mole of sulfur (32 grams) to form 1 mole of zinc sulfide (97 grams) and releases 202.7 kJ of heat.

2. Decomposition

- The prediction of decomposition uses the same source of information, the heat of the formation table.
- If the heat of formation is a high exothermic ($\Delta H$is negative), the compound will be difficult to decompose since this same quantity of energy must be returned to the compound.
- A low heat of formation indicates decomposition would not be difficult

3. Single Replacement

- A prediction can be based on a comparison of the heat of formation of the original compound and that of the compound to be formed

Ex: $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g) \qquad \Delta H = 2 \times -92.3\ kJ$

$2 \times -92.3\ kJ$

$-184.6\ kJ \neq -415.5\ kJ$

This comparison leaves an excess of 230.9 kJ of heat given off, so the reaction would occur.

- Nother simple way to predict single replacement reactions is to check the active positions of the 2 elements in the activity series.
- If the element that is to replace the other in the compound is higher on the chart, the reaction will occur. If it is below, there will be no reaction.

Ex: $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

- In predicting the replacement of hydrogen by zinc in hydrochloric acid, zinc will replace hydrogen and this reaction will occur.
- Most metals in the activity series would replace hydrogen. If a metal such as copper were chosen, no reaction would occur.

4. Double Replacement
- 2 compound seacat where either the cations or anions witch but not both to form 2 new compounds

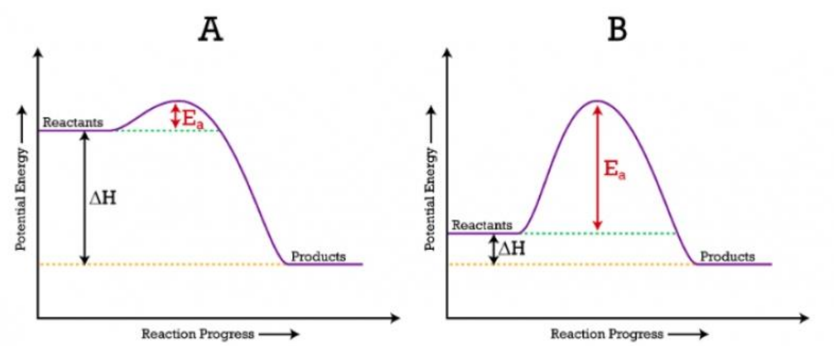
5. Hydrolysis Reactions
- The water and salt react to form an acid and a base

Entropy

- Another driving force to reactions is related to their state of disorder or of randomness called **entropy**.
- A reaction is also driven by a need for a greater degree of disorder.

THERMOCHEMISTRY

- The sum of the internal energy (heat content) of a system is called enthalpy and is symbolized by ΔH.



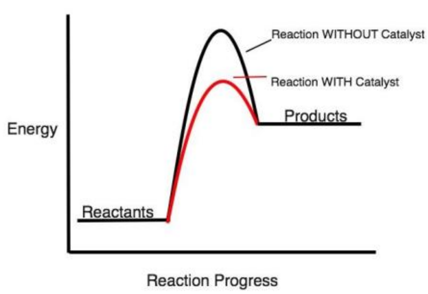
- The heat released by the complete combustion of 1 mole of a substance is called the heat of combustion.
- The First Law of Thermodynamics states that the total energy of the universe is constant and cannot be created or destroyed.

Chapter 9: Rates of Chemical Reactions

FACTORS AFFECTING REACTION RATES

1. **The nature of the reactants.**
- In chemical reactions, some bonds break and others form.
- Therefore, the rates of chemical reactions are affected by the nature of the bonds in the reacting substances.
2. **The surface area exposed.**
- Since most reactions depend on the reactants coming into contact, the surface exposed proportionally affects the rate of the reaction.
3. **The concentration.**
- The reaction rate is usually proportional to the concentrations of the reactants.
- If more molecules or ions of the reactant are in the reaction area, then there is a greater chance that more reactions will occur.
4. **The temperature.**
- A temperature increase of $10^{\circ}C$above room temperature usually causes the reaction rate to double or triple.
    - As temperature increases, the average kinetic energy of the particles involved increases. As a result the particles move faster and have a greater probability of hitting other reactant particles. More energy means more effective collisions, resulting in chemical reactions that form the product.
5. **The presence of a catalyst.**
- Aa catalyst is a substance that increases or decreases the rate of a chemical reactionw/o itself undergoing any permanent chemical change.
- The catalyst provides an alternative pathway by which the reaction can proceed and the activation energy is lower. Thus increasing the rate the reaction comes to completion or equilibrium.

ACTIVATION ENERGY



- A reaction rate may be increased or decreased by affecting the activation energy, the energy needed to cause a reaction to occur.
- A catalyst is a substance introduced into a reaction to speed up the reaction by changing the amount of activation energy needed.

Chapter 10: Chemical Equilibrium

- In some reactions no product is formed to allow the reaction to go to completion; the reactants and products can still interact in both directions.
- $A + B \leftrightarrows C + D$
  - The double arrow indicates that C and D can react to form A and B, while A and B react to form C and D.

EFFECTS OF CHANGING CONDITIONS

Effect of Changing the Concentrations

- When a system at equilibrium is disturbed by adding or removing one of the substances, all the concentrations will change until a new equilibrium point is reached with the same value.
- If the concentration of a reactant in the forward reaction is increased, the equilibrium is shifted to the right, favoring the forward reaction.
- If the concentration of a reactant in the reverse reaction is increased, the equilibrium is displaced to the left.
- Decreases in concentration will produce effects opposite to those produced by increases.

Effect of Temperature on Equilibrium

- If the temperature of a given equilibrium reaction is changed, the reaction will shift to a new equilibrium point.
- If the temperature of a system in equilibrium is raised, the equilibrium is shifted in the direction that absorbs heat.

Effect of Pressure on Equilibrium

- If it is assumed that the total space in which the reaction occurs is constant, the pressure will depend on the total number of molecules in that space.
- An increase in the number of molecules will increase pressure; a decrease in the number of molecules will decrease pressure.
- if the pressure is increased, the reaction that will be favored is the one that will lower pressure or decrease the number of molecules.

Ex: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) + heat\ (at\ equilibrium)$

- If the concentrations of the nitrogen and hydrogen are increased, the forward reaction is increased.
- If the ammonia produced is removed by dissolving it in water, the forward reaction is again favored.
- b/c the reaction is exothermic, the addition of heat must be accounted for.
  - Increasing the temperature causes an increase in molecular motion and collisions thus allowing the product to form more readily.
- An increase in pressure will cause the forward reaction to be favored since the equation shows that four molecules of reactants are forming two molecules of products.

Chapter 11: Acids, Bases, and Salts

Definitions and Properties

Acids

1. **Water (aqueous) solutions of acids conduct electricity.**
- The degree of conduction depends on the acid's degree of ionization.
- A few acids ionize almost completely, while others ionize to only a slight degree.
2. **Acids will react w/ metals that are more active than hydrogen ions to liberate hydrogen.**
- Some acids are also strong oxidizing agents and will not release hydrogen
3. **Acids have the ability to change the color of indicators.**
4. **Acids react w/ bases so that the properties of both are lost to form water and a salt.**
- ACID + BASE → SALT + WATER
5. **Acids react w/ carbonates to release carbon dioxide.**

- The **Arrhenius Theory** states that an acid is a substance that yields hydrogen ions in aqueous solution.
    - Although they are not really not separate ions but become attached to the oxygen of the polar water molecule to form the H_3O^+ion (the hydronium ion).

Ex: The general reaction for the dissociation of an acid, HX, is commonly written as

$$HX \leftrightarrows H^+ + X^-$$

To show the formation of the hydronium ion, the complete equation is:

$$HX + H_2O \leftrightarrows H_3O^+ + X^-$$

Bases

1. **Bases are conductors of electricity in an aqueous solution.**
- Their degrees of conduction depend on their degrees of ionization.
2. **Bases cause a color change in indicators.**
3. **Bases react w/ acids to neutralize each other and form salt and water.**
4. **Bases react w/ fats to form a class of compounds called soaps.**
5. **Aqueous solutions of bases feel slippery, and the stronger bases are very caustic to the skin.**

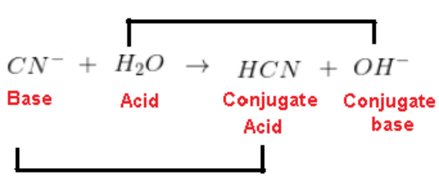
- The **Arrhenius Theory** defines a base as a substance that yields hydroxide ions (OH^-) in an aqueous solution.

- The **Bronsted-Lowry Theory** defines acids as proton donors and bases as proton acceptors.

Ex: $HCl(g) + NH_3(g) \rightarrow NH_4^+(s) + Cl^-(s)$

The HCl is the proton donor or acid, and the ammonia is a Bronsted-Lowry base that accepts the proton.

Conjugate Acids and Bases



- In an acid-base reaction, the original acid gives up its proton to become a conjugate base.
- After losing its proton, the remaining ion is capable of gaining a proton, thus qualifying as a base.
- The original base accepts a proton, so it's now classified as a conjugate acid since it can release this newly acquired proton and behave like an acid.

Acid Concentration Expressed as pH

- pH can be defined as $-log[H^+]$is the concentration of hydrogen ions expressed in moles per liter.
- 100 = 10^2so logarithm of 100, base 10 = 2
- 10,000 = 10^4so logarithm of 10,000, base 10 = 4
- 0.01 =10^{-2}so logarithm of 0.01, base 10 = -2

Ex: Find the pH of a 0.1 molar solution of HCl.
Step 1: Because HCl ionizes almost completely into $H^+$and Cl^-, $[H^+]$ = 0.1 mole/liter.
Step 2:

pH = $-log[H^+]$ so pH = $-log[10^{-1}]$

Step 3: The logarithm of $10^{-1}$is -1 so pH = -(-1)
Step 4: The pH = 1.

- B/c water has a normal H^+concentration of 10^{-7}mole/liter b/c of the slight ionization of water molecules, the water pH is 7 when the water is neither acidic or basic.
- The pOH is the negative logarithm of the hydroxide ion concentration:

  pOH = $-log[OH^-]$

Ex: If the concentration of the hydroxide ion is 10^{-9} M, then the pOH of the solutionis +9.

$[H^+][OH^-]$ =1.0× 10^{-14} AT 298 K

The following relationship can be devised:

pH + pOH = 14.00

- In other words, the sum of the pH and pOH of an aqueous solution at 298 K must always equal 14.00.

TITRATION

- The use of volume measurement in problems to determine the concentrations of “unknown” solutions or solids is called titration.
- A common example uses acid-base reactions.

Ex: If you are given a base of known concentration, 0.10 M NaOH and you want to determine the concentration of an HCl solution.

- First, introduce a measured quantity, 25,0 mL of the NaOH into a flask by using a pipette.
- Next, introduce 2 drops of a suitable indicator. b/c NaOH and HCl are considered a strong base and acid, an indicator that changes color in the middle pH range would be appropriate. Litmus solution could be used as its blue in basic solution but changes to red in acidic.
- Slowly introduce HCl until the color change occurs, this is called the **end point**.
- The point at which enough acid is added to neutralize all the standard solution in the flask is called the **equivalence point**.

SALTS

- A salt is an ionic compound containing positive ions other than hydrogen ions and negative ions other than hydroxide ions.
- Five methods of preparing a salt are:

1. Neutralization reaction.
- $2HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$
- ACID + BASE → SALT + WATER
2. Single replacement reaction.
- An active metal replaces hydrogen in an acid.
- $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
3. Direct combination of elements.
- $Fe(s) + S(s) \rightarrow FeS(s)$
4. Double Replacement

- When solutions of 2 soluble salts are mixed, they form an insoluble salt compound
- $AgNO_3(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + AgCl(s)$

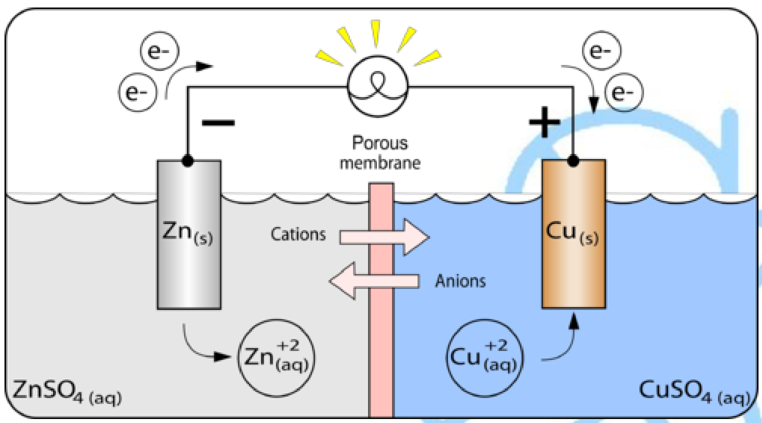
5. Reaction of a metallic oxide w/ a nonmetallic oxide.

- $MgO(s) + SiO_2(s) \rightarrow MgSiO_3(s)$

Chapter 12: Oxidation-Reduction and Electrochemistry

- The branch of chemistry that deals w/ electricity-related applications of oxidation-reduction is called electrochemistry.
- Oxidation-reduction reactions involve a transfer of electrons from the substance oxidized to the substance reduced.

Voltaic Cells (Galvanic Cells)



- The substance that is oxidized during the reaction is separated from the substance that is reduced during the reaction.
- The electron transfer is accompanied by a transfer of electrical energy instead of heat
- One means separating oxidation and reduction half-reactions is w/ a porous barrier, which prevents the metal atoms of one-half reaction from mixing w/ the ions of the other half-reaction.
- Ions can move through the porous barrier.
- Electrons can be transferred from one side to the other through an external connecting wire.
- Electric current moves in a closed loop path so this movement of electrons through the wire is balanced by the movement of ions in the solution.
- The zinc strip is in an aqueous solution of $ZnSO_4$; the copper strip, in an aqueous solution of $CuSO_4$.
- An electrode is a conductor used to establish electrical contact w/ a nonmetallic part of the circuit.
- A single electrode immersed in a solution of its ions is a **half-cell**.
- The zinc strp in aqueous $ZnSO_4$is an **anode**, the electrode where oxidation takes place.
- The copper strip in $CuSO_4$is a **cathode**, the electrode where reduction takes place.
- The copper half-cell can be written as Cu^{2+}/Cu, and the zinc half-cell as Zn^{2+}/Zn.
- The two half-cells together constitute an electrochemical cell. An **electrochemical cel**l is a system of electrodes and electrolytes in which either chemical reactions produce electrical energy or an electric current produces chemical change.

Half-reactions: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad$ $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
Oxidized Half-reaction: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-2}$
Reduction Half-reaction: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

- The gain of electrons is **reduction** and the loss of electrons is **oxidation**.
- “LEO the lion says GER”: “Loss of Electrons is Oxidation” and “Gain of Electrons is Reduction”
- The fact that the zinc in this reaction is oxidized by giving up electrons makes it possible for electrons to be gained by the copper, which is acting as a **reducing agent**.

- Because the copper in this reaction is being reduced by gaining electrons, electrons can be lost by the zinc, which is acting as an **oxidizing agent**.

Electrolytic Cells

- The redox reactions do not occur spontaneously but can be forced to take place by supplying energy w/ an external current, these are called **electrolytic reactions**.
- In electroplating, where electrolysis is used to coat a material w/ a layer of metal, the object to be plated is made the cathode in the reaction.

BALANCING REDOX EQUATIONS

The Rules for Assigning an Oxidation State

- NOTE: The sum of the oxidation states must equal zero for an electrically neutral compound. For an ion, the sum of the oxidation states must equal the charge on theion, including polyatomic ions.

1. The oxidation state of an atom in an element is zero.
2. The oxidation state of a monatomic ion is the same as its charge.
3. The oxidation state of fluorine is -1 in it's compounds.
4. The oxidation state of oxygen is usually -2 and its compounds.
    a. One exception to this rule occurs when oxygen is bonded to flooring in the oxidation state of chlorine takes precedence.
    b. Another exception occurs in peroxide compounds, where the oxidation state is assigned the value of -1.
5. The oxidation state of hydrogen in most compounds is +1.
    a. In hydrides, there is an exception. Oxygen is assigned the value of -1.

The Electron Shift Method

- The oxidation state method breaks the reaction into two half reactions, the reduction reaction and the oxidation reaction.

Ex: Balance the equation $HCl(aq) + MnO_2(s) \rightarrow H_2O(l) + MnCl_2(aq) + Cl_2(g)$

1. The oxidation states are written above all the elements in the equation .

+1 -1 +4 -2 +1 -2 +2 -1 0

$$HCl(aq) + MnO_2(s) \rightarrow H_2O(l) + MnCl_2(aq) + Cl_2(g)$$

2. The oxidation state of the manganese has changed from +4 to +2, while the oxidation state of the chlorine that emerges as a gas has changed from -1 to 0. A change from +4 to +2 is a reduction in the oxidation state and a game in electrons. These are written as the half reactions of the ions involved with the electron changes.

NOTE: oxidation states have a + or - on the left of the number and that ionic charges are shown on the right of the number.

$Mn^{4+}(aq) + 2e^- \rightarrow Mn^{2+}(aq)$ Reduction reaction
$Cl^{1-}(aq) \rightarrow Cl(g) + 1e^-$ Oxidation reaction

3. The next step is to assure that the number of atoms required by the formulas in the reactants equals the number of atoms in the products. Notice that the product Cl_2 in the original equation is a diatomic molecule and requires 2 atoms in the molecular formula. This means that there must be 2 Cl^- ions provided on the left side of the equation. A total of $2e^-$ are lost, one for each Cl^-.

$Mn^{4+}(aq) + 2e^- \rightarrow Mn^{2+}(aq)$ Reduction reaction
$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$ Oxidation reaction

4. If the number of electrons gained in reduction did not equal the number electrons lost in oxidation, we would have to multiply either one reaction line or both by an appropriate number to ensure that the number of electrons lost equaled the number of electrons gained.
5. Next add the two half reactions.

$$Mn^{4+}(aq) + 2Cl^-(aq) \rightarrow Mn^{2+} + Cl_2(g)$$

6. Now go back to the original equation and insert the coefficients.
   $2HCl(aq) + 1MnO_2(s) \rightarrow ?\, H_2O(l) + 1MnCl_2(aq) + Cl_2(g)$
7. From the numbers thus established, the remaining coefficients can be deduced by infection. Notice that 2 more molecules of HCl are required to furnish the chlorine for the $MnCl_2$, and the 2 atoms of oxygen in MnO_2form $2H_2O$.
8. The final balanced equation is:

$$4HCl(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$$

Sources of photos:

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