Chemistry at Argo Community High School

Updated 2022

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Chemistry Curriculum Map
Semester One

Unit 1. Atoms and Elements

Number of class periods – 10
Number of Labs – 1

In this unit students are introduced to the atom and its constituent parts. Determining the number of protons and electrons by using the periodic table is an important skill that students will use throughout the entire curriculum. The group names, metal nonmetal location, and group and period numbers are equally important and used throughout the semester.

Topic 1.1 Elements and Symbols

1.1.1 Enduring Understanding

The elements listed on the periodic table make up all matter in the known universe.

1.1.1.1 Learning Objective

The periodic table consists of elements that have symbols and names.

1.1.1.1.1 Essential Knowledge

Elements are represented with either a single capital letter or a capital letter followed by a lowercase letter.

Knowing the element names and symbols is essential to understanding chemistry.

1.1.1.1.2 Exclusion Statement

The elements expected to be memorized are listed in **Error! Reference source not found.**. Element quizzes do not begin until Unit 5.

Topic 1.2 The Periodic Table

1.2.1 Enduring Understanding

The periodic table is arranged to display information about the elements.

1.2.1.1 Learning Objective

The periodic table has periods and groups. Some of those groups have names.

1.2.1.1.1 Essential Knowledge

On every periodic table, periods run horizontally and groups run vertically. Periods are numbered 1 – 7 and Groups are numbered 1 – 18.

The periodic table can be divided into metals and nonmetals with metalloids dividing the two.

Groups 1,2 and 13 – 18 are the representative elements.

Groups 3 – 12 are the transition elements.

Certain groups have names that must be memorized. Group 1 = alkali metals, group 2 = alkaline earth metals, group 17 = halogens, group 18 = noble gases

Topic 1.3 The Atom

1.3.1 Enduring Understanding
The atom is the smallest particle of an element that retains the properties of that element.

1.3.1.1 Learning Objective
The atom is made of protons, neutrons, and electrons.

1.3.1.1.1 Essential Knowledge
Each element has a unique number of protons.

Atoms are neutral. Because of this, the number of protons and electrons in an atom are the same.

The atomic number is the number of protons in an element. The mass number is the sum of the protons and neutrons. Electrons are relatively massless.

Atoms of the same element can have different numbers of neutrons which leads to different isotopic masses. Two atoms of the same element with different numbers of neutrons are called isotopes.

The mass spectrum of a sample containing a single element can be used to determine the identity of the isotopes of that element and the relative abundance of each isotope in nature.

The average atomic mass of an element can be estimated from the weighted average of the isotopic masses using the mass of each isotope and its relative abundance.

1.3.1.1.2 Exclusion Statement
The history of the atom and atomic models will not be assessed.

Unit 2. Measurement

Number of class periods – 15
Number of Labs – 3

Topic 2.1 Units of Measurement

2.1.1 Learning Objective

A measurement should contain an actual and estimated value according to the device being used.

2.1.1.1 Essential Knowledge

A measurement includes a number and a unit.

A measured value should be recorded one place smaller than the smallest measurement on the device being used to measure.

Accuracy of a measurement is determined by percent error and precision of a measurement is determined by standard deviation.

2.1.1.2 Exclusion Statement

Uncertainty will not be assessed. Students are not expected to calculate standard deviation, just interpret the standard deviation of groups of measurements.

2.1.2 Learning Objective

Significant figures are used to record a calculated value to the correct number of digits.

2.1.2.1 Essential Knowledge

Some numbers are significant and some numbers are not significant. All nonzero numbers are significant. Zeros between nonzero numbers are significant. All other zeros are not significant.

The fewest number of significant figures in a question should be used as the number of significant figures in an answer.

2.1.2.2 Exclusion Statement

The rules for significant figures for addition and subtraction will not be used or assessed. The purpose of significant figures in this course are to round calculations properly. The determination of the number of significant figures in a number will not be assessed.

Topic 2.2 Scientific Notation

2.2.1 Learning Objective

Use scientific notation to represent very large and very small numbers.

2.2.1.1 Essential Knowledge

Very large and very small numbers are represented with scientific notation. Scientific notation consists of a real number (called a mantissa and is usually between 1 and 10) times ten raised to the power of an exponent.

Positive exponents are used to represent numbers larger than 1. Negative exponents are used to represent numbers smaller than 1.

The exponent indicates the number of times a decimal has been moved to place the number in scientific notation. Positive exponents have the decimal moved to the right to convert to a standard number. Negative exponents have the decimal moved to the left to convert to a standard number.

2.2.2 Learning Objective

Use a scientific calculator to perform calculations with scientific notation.

2.2.2.1 Essential Knowledge

On the school issued TI-30XS, the button “ $x10^n$ ” is used for scientific notation.

2.2.2.2 Exclusion Statement

Scientific Notation calculations without a calculator will not be assessed.

Topic 2.3 Density

2.3.1 Learning Objective

Density is a physical property of a substance and can be used to identify the substance.

2.3.1.1 Essential Knowledge

Density can be calculated with the formula:

$$D = \frac{m}{V}$$

where:

D = density, measured in g/mL or g/cm^3
m = mass, measured in g
V = volume, measured in cm^3 of mL

The density of a substance can be used to determine mass or volume if the other variable is known.

Mass is typically directly measured on an electronic balance.

Volume can be measured by direct measurement or indirect measurement. The volume of a geometrical shape can be measured and calculated directly, e.g. the volume of a cube is $V = s^3$, where s is the length of the side. The volume of a liquid is measured with a graduated cylinder. The volume of an irregular shaped object is measured by volume displacement using a graduated cylinder.

Unit 3. Matter and Energy

Number of class periods – 15
Number of Labs – 3

Topic 3.1 Classification of Matter

3.1.1 Learning Objective

Matter can be classified as a pure substance or a mixture.

3.1.1.1 Essential Knowledge

A pure substance is either a compound or an element. A compound is made of two or more atoms chemically bonded while an element is made of only one type of atom.

Compounds can be broken down chemically into smaller constituents while elements cannot be broken down into smaller constituents.

A pure substance always has the same composition by mass while a mixture does not.

A mixture contains compounds or elements of two or more types and can be homogeneous or heterogeneous.

Particle pictures can be used to represent compounds, elements, and mixtures.

Mixtures can be separated physically. One method of separating a mixture is chromatography.

3.1.1.2 Exclusion Statements

Chromatography will be discussed qualitatively only.

Topic 3.2 Energy

3.2.1 Learning Objective

All energy can be classified as kinetic or potential energy.

3.2.1.1 Essential Knowledge

In Chemistry, kinetic energy is defined as the movement of the particles. The average kinetic energy is related to the temperature of the substance. Two substances at the same temperature have the same average kinetic energy. An increase in temperature is accompanied by an increase in average kinetic energy. A decrease in temperature is accompanied by a decrease in average kinetic energy.

In Chemistry, potential energy is defined as the energy as a result of a force of attraction or repulsion. Forces of attraction include ionic bonds (Topic 5.1 below), covalent bonds (Topic 9.1 below), and intermolecular forces (Topic 9.4 below). A change in only potential energy is NOT accompanied by a change in temperature.

Energy can be measured in Joules or calories. Dimensional analysis can be used to convert between Joules, calories, kilojoules, and kilocalories.

Topic 3.3 Specific Heat Capacity

3.3.1 Learning Objective
Specific Heat Capacity, C_s, is a physical property of a substance and can be used to identify the substance.

3.3.1.1 Essential Knowledge
Each substance has its own specific heat capacity. The specific heat capacity of a substance can be calculated as shown below:

$$C_s = \frac{q}{m\Delta T}$$

where:

C_s = specific heat capacity, measured in J/(g °C)
q = heat energy, measured in J
m = mass, measured in g
ΔT = change in temperature, measured in °C or K

The greater the specific heat capacity of a substance the more energy is required to change the temperature.

3.3.2 Learning Objective
Substances will transfer energy through direct contact and will do so until thermal equilibrium has been reached. Energy is transferred from the high energy substance to the low energy substance.

3.3.2.1 Essential Knowledge
Substances transfer energy when they are in direct contact. That transfer of energy will continue until both substances reach thermal equilibrium resulting in both substances having the same final temperature.

Energy is always transferred from the high energy substance to the low energy substance.

The amount of energy lost by the high energy substance is equal in magnitude but opposite in sign to the amount of energy gained by the low energy substance.

Unit 4. Electronic Structure and Periodic Trends

Number of class periods – 15
Number of Labs – 2

Topic 4.1 Atomic Spectra and Energy Levels

4.1.1 Learning Objective

Photons of light have finite energy that can be defined by their wavelength and frequency.

4.1.1.1 Essential Knowledge

A photon is a quantum of light energy. A photon's energy is inversely proportional to the wavelength or and directly proportional to the frequency of the photon, as shown in the equations below:

$$E = h\nu = \frac{hc}{\lambda}$$

where:

E = energy, measured in J
h = Planck's constant, 6.626×10^{-34} Js
ν = frequency, measured in Hz
c = speed of light, 3.00×10^{8} m/s or 3×10^{17} nm/s
λ = wavelength, measured in m or nm

The photons energy falls somewhere within the electromagnetic spectrum, of which visible light is a small portion.

4.1.2 Learning Objective

Every element gives off a distinct frequency of light and that frequency can be used to identify the element.

4.1.2.1 Essential Knowledge

Elements have specific number of protons that attract the electrons in the atom. That attraction determines the placement of electrons within the atom. The placement of electrons within the atom determines their energy according to Coulomb's Law. When atoms absorb energy the electrons can move further away from the nucleus and are in the "excited" state. Eventually that same electron releases the energy absorbed and falls back to its original energy "ground" state. The energy given off by the atom creates a line spectra which can be used to identify an element.

Topic 4.2 Electron Configuration and Orbital Notation

4.2.1 Learning Objective

The electron configuration of an element can be used to determine properties of that element.

4.2.1.1 Essential Knowledge

In atoms, the electrons can be thought of being in "shells (energy levels)" and "subshells (sublevels)", as described by the electron configuration. Inner electrons are called core electrons, and outer electrons are called valence electrons. The electron configuration is explained by quantum mechanics, as delineated by the Aufbau principle and exemplified in the periodic table of the elements.

The periodic table can be used to determine the electron configuration of an element. The periods correspond to the energy level of the electron and the table can be divided into s-block, p-block, and d-block elements. The Aufbau principle and Hund's rule are used to place electrons in orbitals. Orbital notation is shorthand notation for electron configuration and consists of boxes and arrows pointing up and down. The boxes represent the suborbital and the arrows represent electrons and their spin.

4.2.1.2 Exclusion Statements

Exceptions to normal electron configuration (e.g. Cu) will not be assessed.
Electron configuration beyond the d-block will not be assessed.
Students will be given completed orbital notations for elements up to the d-block therefore memorizing the number of suborbital boxes needed is unnecessary.

Topic 4.3 Periodic Trends

4.3.1 Learning Objective

Explain the relationship between trends in atomic properties of elements and electronic structure and periodicity.

4.3.1.1 Essential Knowledge

The organization of the periodic table is based on the recurring properties of the elements and explained by the pattern of electron configurations.

Trends in atomic properties within the periodic table can be qualitatively understood through the position of the element on the periodic table, Coulomb's law, the shell model, and the concept of shielding/effective nuclear charge. These properties include valence electrons, atomic radius, ionization energy, and successive ionization energy. (Ionic radii and electronegativity are taught later.)

The relative energy required to remove an electron from different subshells of an atom or ion or from the same subshell of different atoms or ions (ionization energy) can be estimated through a qualitative application of Coulomb's law. This energy is relative to the distance from the nucleus and the effective (shield) charge of the nucleus.

Unit 5. Ionic and Molecular Compounds

Number of class periods – 15
Number of Labs –

Topic 5.1 Ions: Transfer of Electrons

5.1.1 Learning Objective
In an ionic bond, metals lose electrons and nonmetals gain electrons in order to become stable.

5.1.1.1 Essential Knowledge
Metals will lose electrons in part due to their relatively large atomic radius and low ionization energy. When metals lose an electron they become positively charged. The number of electrons lost is dependent upon valence electrons and their location on the periodic table. Some transition metals have two or more different cations (variable charge) and Roman numerals are used to identify the specific transition metal.

Nonmetals will gain electrons in part due to the relatively small atomic radius and high ionization energy. When nonmetals gain an electron they become negatively charged. The number of electrons gained is dependent upon valence electrons and their location on the periodic table.

5.1.1.2 Exclusion Statement
Students do not have to memorize the charges of transition metals.
Students do not need to memorize names and formulas of polyatomic ions.

Topic 5.2 Naming and Writing Formulas for Ionic Compounds

5.2.1 Learning Objective
Cations and anions will form an ionic bond with an overall neutral charge.

5.2.1.1 Essential Knowledge
Molecules have a neutral charge. In order to be neutral the entire positive charge of the cations must cancel with the entire negative charge of the anions. Subscripts are used to indicate the number of cations and anions present in a molecular formula. When more than one polyatomic ion is required the ion is placed in parentheses and the number of ions required is given as a subscript.

5.2.2 Learning Objective
Ionic compounds are named according to their ions.

5.2.2.1 Essential Knowledge
Monatomic cations are named after the metal, i.e. Na^+ is named “sodium ion”. Monatomic anions are named after the nonmetal and end in “-ide”, i.e., O^{2-} is named “oxide”. Binary ionic compounds combine the name of the cation and the name of the anion, i.e. Na_2O is named “sodium oxide”. The names of polyatomic ions are not changed, i.e. Na^+ and NO_3^- is named “sodium nitrate”.

5.2.2.2 Exclusion Statement
Students do not need to name ionic compounds that contain a transition metal, i.e. students will not need to go from CuO to “copper(II) oxide”. However, students will need to write the formula from the name, i.e. “copper(II) oxide” converted to CuO.

Topic 5.3 Properties of Ionic Bonds

5.3.1 Learning Objective
Ionic compounds have specific properties and those properties are different from other types of bonds.

5.3.1.1 Essential Knowledge
Ionic bonds have a high melting point, are generally soluble in water, and can conduct electricity when dissolved in water. Covalent bonds (Topic 9.1 below) typically have lower melting points, are mostly insoluble in water, and can not conduct electricity in water.

5.3.1.2 Exclusion Statement

Electrolyte terminology will not be assessed in this unit. Electrolytes will be discussed in Unit 11. Covalent bonds and how covalent bonds form will not be assessed in this unit. Covalent bonds will be discussed in 0.

Chemistry Curriculum Map
Semester Two

Unit 6. Chemical Quantities

Number of class periods – 15
Number of Labs –

Topic 6.1 The Mole and Molar Mass

6.1.1 Learning Objective

The mole is an amount of a substance used in science.

6.1.1.1 Essential Knowledge

Avogadro's number ($N_A = 6.022 \times 10^{23}$) provides the connection between the number of moles in a pure sample of a substance and the number of constituent particles (or formula units) of that substance. One cannot count particles directly while performing laboratory work. Thus, there must be a connection between the masses of substances reacting and the actual number of particles undergoing chemical changes.

6.1.1.2 Exclusion Statement

Calculations involving Avogadro's number will not be assessed.

6.1.2 Learning Objective

A mole of a substance has a mass that can be determined by using the periodic table.

6.1.2.1 Essential Knowledge

Expressing the mass of an individual atom or molecule will always be numerically equal to the molar mass of that substance in grams. Thus, there is a quantitative connection between the mass of a substance and the number of particles that the substance contains. To determine the mass of one mole of a substance, termed molar mass, the mass of the individual atoms that make up the substance are added together.

$$Molar\ mass = \frac{Grams}{Moles}\ or\ g/mol$$

Topic 6.2 Calculations Using Molar Mass

6.2.1 Learning Objective

Molar mass can be used to convert from grams to moles or moles to grams of a substance.

6.2.1.1 Essential Knowledge

The molar mass of a substance is the number of grams in one mole of that substance. Molar mass can be used to calculate the amount (moles) or mass (grams) for more or less than one mole of a substance. To convert from grams of a substance to moles the grams are divided by the molar mass. To convert from moles of a substance to grams the moles are multiplied by the molar mass. Dimensional analysis, i.e. T-charts, must be used to convert between grams and moles.

Topic 6.3 Mass Percent Composition

6.3.1 Learning Objective

The percent by mass of a substance can be used to determine the mass of an element in any given sample of a compound.

6.3.1.1 Essential Knowledge

The mass percent composition of an element in a compound is calculated by dividing the product of the molar mass of the element and the number of times that element is found in the substance by the molar mass of the substance and finally multiplying by 100.

$$\%\ by\ Mass\ of\ Element = \frac{(Molar\ Mass\ of\ Element)\ x\ (Number\ of\ times\ Element\ is\ Found\ in\ Substance)}{Molar\ mass\ of\ substance}\ x\ 100$$

The mass percent composition can be used to determine the mass of the element in any amount of the substance. To do so, multiply the percentage by mass of the element by the mass of substance present.

$$Mass\ of\ element\ in\ sample = \%\ by\ mass\ of\ element\ x\ mass\ of\ substance$$

OPTIONAL

Topic 6.4 Empirical and Molecular Formulas

6.4.1 Learning Objective

The empirical formula of a compound is the smallest mole ratio of elements for a substance. The molecular formula is the actual mole ratio of elements for a substance.

6.4.1.1 Essential Knowledge

The empirical formula is the smallest mole ratio of elements in a substance. An empirical formula can not be further divided into whole numbers of moles of each element. An empirical formula can be calculated from the percent composition by mass of a compound by first converting to grams, then moles, and finally finding the smallest whole number ratio of moles for each element. A molecular formula can be determined by comparing the molar mass of the molecule to the molar mass of the empirical formula.

Unit 7. Chemical Reactions

Topic 7.1 Equations for Chemical Reactions

7.1.1 Learning Objective

Chemical reactions contain symbols that relay information about the reaction.

7.1.1.1 Essential Knowledge

A chemical reaction gives the reactants and products involved in a chemical reaction. Those reactants and products are given phase symbols and are separated by some form of a yield symbol.

7.1.2 Learning Objective

Some signs of a chemical reaction may include a change in energy in the form of heat, giving off light, giving off a gas, or a change in color.

7.1.2.1 Essential Knowledge

There are four signs of a chemical reaction: change in energy, giving off light, giving off a gas, or a change in color. A change in energy may absorb or release energy and can be measured with a thermometer. A reactant may also give off light as a sign of a chemical change. If bubbles form during a chemical reaction then a gas is considered to be given off. A permanent change in color may also indicate a chemical reaction.

Topic 7.2 Balancing a Chemical Equation

7.2.1 Learning Objective

Because of the Law of Conservation of Mass, a balanced chemical equation has the same number and types of atoms on both sides of the yields symbol.

7.2.1.1 Essential Knowledge

The Law of Conservation of Mass states that mass cannot be created or destroyed in a normal chemical reaction. Because of this, a chemical equation must be balanced by making the number and types of atoms in the reactants the same as the number and types of atoms in the products. To balance a chemical equation, coefficients, which are numbers in front of the particle, are used. The subscripts in the chemical equation cannot be changed.

Topic 7.3 Types of Reactions

7.3.1 Learning Objective

Many chemical reactions can be sorted into one of five types; single replacement, double replacement, synthesis, decomposition, and combustion.

7.3.1.1 Essential Knowledge

There are five types of chemical reactions; single replacement, double replacement, synthesis, decomposition, and combustion. A single replacement reaction has a compound that reacts with an element. A double replacement reaction has two ionic compounds that react. A synthesis reaction combines two or more reactants into one product. A decomposition reaction has one reactant that becomes two or more products. A combustion reaction is strictly defined as an organic hydrocarbon reacting with oxygen gas, O_2, to form water and carbon dioxide gas.

7.3.2 Learning Objective

Predict the products of a single replacement and double replacement reaction. In a double replacement reaction predict the formation of a precipitate.

7.3.2.1 Essential Knowledge

In a single replacement reaction a metal will replace a metal and a nonmetal will replace a nonmetal. In a double replacement reaction the cation of reactant 1 will combine with the anion of reactant 2 while the anion of reactant 1 will combine with the cation of reactant 2. In both single and double replacement reactions the atoms/ions will combine such that the overall charge of the particle is neutral. For double replacement reactions, a precipitate may form as a product. Using a solubility chart, the precipitate can be determined from the combination of cations and anions.

7.3.2.2 Exclusion Statement

Students are not expected to predict the formula of the product. Reaction prediction will be assessed with multiple choice questions only and not free response.

Unit 8. Chemical Quantities in Reactions

Topic 8.1 Mole Relationships in Chemical Equations

8.1.1 Learning Objective

When an equation is properly balanced, the coefficients can be used to give a ratio of moles between the reactants and products.

8.1.1.1 Essential Knowledge

The coefficients of a balanced chemical equation gives the mole ratio between reactants and products. That mole ratio can be used to convert from moles of one species to moles of a different species using a t-chart.

Topic 8.2 Mass Calculations for Reactions

8.2.1 Learning Objective

The mass of reactants and products used/formed can be predicted through the use of stoichiometry with a balanced chemical reaction.

8.2.1.1 Essential Knowledge

The mass of a reactant or product can be converted to moles by using the molar mass (see Topic 6.2 above, on page 15). By using the mole ratio from Topic 8.1 above the moles of a different reactant or product can be determined and converted to grams.

Topic 8.3 Limiting Reactants

8.3.1 Learning Objective

The mass of reactants present will determine the amount of products formed. Sometimes one reactant is completely consumed before all others.

8.3.1.1 Essential Knowledge

A limiting reactant limits how much product is formed. The limiting reactant is completely consumed in a chemical reaction while a portion of the excess reactant remains unreacted upon reaction completion. To determine a limiting reactant the masses of all reactants are converted via stoichiometry to the mass of a single product. Whichever reactant yields the smallest mass of product is the limiting reactant and all other reactants are termed the excess reactant.

8.3.1.2 Exclusion Statement

The amount of excess reactant remaining will not be assessed.

OPTIONAL

Topic 8.4 Percent Yield

OPTIONAL

Topic 8.5 Energy in Chemical Reactions

8.5.1 Learning Objective

The energy in a chemical reaction, either exothermic or endothermic, can be used in stoichiometry similar to mass or moles.

8.5.1.1 Essential Knowledge

A chemical reaction may absorb or release energy. Absorbing energy is termed endothermic and has a positive change in energy while releasing energy is termed exothermic and has a negative change in energy. The energy released or absorbed depends on the amount of reactants present or products formed. The energy in a chemical reaction can be determined with stoichiometry.

Unit 9. Properties of Solids and Liquids

Topic 9.1 Covalent Bonds, Electron-dot Formulas, and Lewis Structures

9.1.1 Learning Objective

The Lewis structure shows how atoms are bonded to form a particle.

9.1.1.1 Essential Knowledge

Valence electrons of single atoms can be indicated with electron-dot formulas, where each dot represents a valence electron. Covalent bonds occur when atoms share electrons to become more stable. Most atoms, except hydrogen, will share electrons in order to acquire a full s- and p-orbital with eight electrons. Hydrogen will share electrons in order to acquire two electrons. When atoms share electrons it is termed a covalent bond and can be two electrons (single covalent bond), four electrons (double covalent bond) or six electrons (triple covalent bond). Atoms may also have lone pair electrons that are not shared when bonding.

9.1.2 Learning Objective

Lewis structures can be drawn for covalently bonded particles.

9.1.2.1 Essential Knowledge

The steps for drawing a Lewis structure are:

1) Determine the total number of valence electrons.
2) Determine the central atom. If one atom is not obvious then pick the least electronegative atom.
3) Place all other atoms around the central atom. Connect the atoms to the central atom by using a single bond.
4) Give all atoms 8 total electrons except hydrogen.
5) Compare the electrons on the Lewis structure with the number of valence electrons from step 1.
    a. If the electrons match then the Lewis structure is complete.
    b. If too many electrons are shown then make double or triple bonds.

9.1.2.2 Exclusion Statement

Particles that exceed the octet rule will not be assessed. Resonance structures will not be assessed.

Topic 9.2 Shapes of Molecules and Ions (VSEPR Theory)

9.2.1 Learning Objective

VSEPR theory determines the shape, or molecular geometry, of a molecule or polyatomic ion.

9.2.1.1 Essential Knowledge

The repulsion of the atoms in the Lewis structure of a molecule/polyatomic ion determines the molecular geometry. Lone pair electrons repel with greater force than bonded atoms. Possible molecular geometries include linear, trigonal planar, bent (120º), tetrahedral, trigonal pyramid, and bent (109.5º).

9.2.1.2 Exclusion Statement

Molecular geometries do not need to be memorized; students are allowed to use the table of molecular geometries on the “Pink” sheet. Trigonal bipyramidal and octahedral geometries will not be assessed.

Topic 9.3 Electronegativity and Polarity

9.3.1 Learning Objective

Some bonds result in an unequal sharing of electrons.

9.3.1.1 Essential Knowledge

Electronegativity is a measure of an atoms ability to attract electrons while covalently bonded. Electronegativity is a periodic trend that increases across a period and decreases down a group. A difference in electronegativity results in a polar bond while a very small or no difference in electronegativity results in a nonpolar bond. A bond between two of the same element and a bond between C and H is considered nonpolar. All other covalent bonds should be considered polar. A polar bond has a positive end and a negative end and is referred to as a “dipole”. The more electronegative atom attracts the shared electrons with greater force and is the negative end of the polar bond while the less electronegative atom is the positive end of the polar bond.

9.3.2 Learning Objective
The molecular geometry and whether polar bonds are present determine the overall polarity of a molecule.

9.3.2.1 Essential Knowledge
A molecule made of nonpolar bonds will be nonpolar while a molecule made of polar bonds may be polar. If the dipoles of a molecule cancel out due to the molecular geometry then the molecule is nonpolar. If the dipoles do not cancel out then the molecule is polar.

Topic 9.4 Attractive Forces in Compounds

9.4.1 Learning Objective
Intermolecular forces arise from forces of attraction between different molecules and have varying strengths that is dependent upon the makeup of the molecule.

9.4.1.1 Essential Knowledge
Intermolecular forces (IMF) result from the attraction between two different molecules. London dispersion forces are found in nonpolar and polar molecules and are the result of a temporary dipole as a result of the movement of electrons. Dipole-dipole forces are found in polar molecules only and are the result of attractions between a positive end and negative end between two permanent dipoles. Hydrogen bonding forces are exhibited by molecules with a H bound to the very electronegative N, O, or F resulting in a force of attraction with a different molecules O, N, or F. Generally, the more IMF present the stronger the force of attraction. If similar IMF are present then the molecule with a greater number of electrons will generally have a greater IMF.

Topic 9.5 Changes of State

9.5.1 Learning Objective
A change in the strength of IMF is associated with a change in state.

9.5.1.1 Essential Knowledge
When a single substance is considered, the solid has the strongest IMF, followed by the liquid, and finally the gaseous state. To change from solid to liquid to gas energy must be absorbed in order to weaken or break the IMF. A heating curve displays how energy is used to convert between states of matter. A heating curve consists of portions with a change in temperature and portions with no change in temperature. When the temperature changes the motion of the particles change. When the temperature does not change the strength of the IMF of the particles change.

9.5.2 Learning Objective
A substance with stronger IMF will have a greater melting and boiling point.

9.5.2.1 Essential Knowledge
Melting and boiling are physical changes. When a substance melts or boils the IMF are weakened or broken but the molecule itself does not change. The greater the IMF present the stronger the force of attraction between molecules. The stronger the force of attraction between molecule the more energy required to separate the molecules. The more energy required to separate the molecules the higher the melting and boiling point.

Unit 10. Gases

Topic 10.1 Properties of Gases and Gas Pressure

10.1.1 Learning Objective

Gases can be described with the physical properties of pressure, volume, temperature, and amount.

10.1.1.1 Essential Knowledge

The pressure, volume, temperature, and amount of a gas can be used to describe the physical properties of a gas. The pressure of a gas arises from the collision of particles with the side of the container. Pressure can be measured in atmospheres (atm), kiloPascals (kPa), and millimeters of mercury (mm Hg). It is often necessary to convert between the units of pressure. Atmospheric pressure describes the pressure exerted by the molecules of the atmosphere. The volume of a gas is typically measured in liters. A gas will occupy the entire volume of the container. The temperature of a gas must be measured in Kelvin. The amount of a gas is measured in moles.

$$1\ atm = 101.3\ kPa = 760\ torr$$

$$T_K = T_C + 273$$

where:

T_K = temperature measured in Kelvin
T_C = temperature measured in Celsius

10.1.2 Learning Objective

The kinetic molecular theory is used to explain the behavior of a gas.

10.1.2.1 Essential Knowledge

The kinetic molecular theory has five postulates that can be used to explain the behavior of a gas.

1. Gases consist of small particles with high velocities
2. The attractive forces between gas particles is negligible
3. The volume occupied by the gas particle is negligible
4. Gas particles are in constant, straight path motion
5. The average kinetic energy of gas molecules is proportional to the Kelvin temperature

OPTIONAL

Topic 10.2 The Combined Gas Law

10.2.1 Learning Objective

The Combined Gas Law can be used to describe the properties of a gas that is undergoing a change in volume, temperature, or pressure.

10.2.1.1 Essential Knowledge

The Combined Gas Law can be used to describe the conditions of a gas that is undergoing a change in volume, temperature, or pressure. While the volume and pressure can be measured in any units the temperature must be measured in Kelvin. In the Combined Gas Law a subscript of 1 or 2 is used to denote the pressure, volume, and temperature that occurs at the same time. The Combined Gas Law is:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

where:

P_1, V_1, and T_1 = the pressure, volume, and temperature at time "1"
P_2, V_2, and T_2 = the pressure, volume, and temperature at time "2"

STP stands for standard temperature and pressure and is 1 atm and 0 ℃.

10.2.1.2 Exclusion Statement

Students do not need to memorize which equation is Boyle's law, Charles's law, or Gay-Lussac's law.

Topic 10.3 The Ideal Gas Law

10.3.1 Learning Objective
The Ideal Gas Law is used to completely describe the physical properties of a gas that is not undergoing any change in conditions.

10.3.1.1 Essential Knowledge
A gas can be described by using the Ideal Gas Law.

$$PV = nRT$$

where:

P = pressure (atm)
V = volume (L)
n = moles (mols)
R = gas constant (0.0821 $\frac{L \cdot atm}{mol \cdot K}$)
T = temperature (K)

Grams and molar mass can be substituted into the Ideal Gas Law.

$$PV = \frac{gRT}{MM}$$

where:

P = pressure (atm)
V = volume (L)
g = mass (grams)
R = gas constant (0.0821 $\frac{L \cdot atm}{mol \cdot K}$)
T = temperature (K)
MM = molar mass (g/mol)

OPTIONAL

Topic 10.4 Partial Pressure (Dalton's Law)

10.4.1 Learning Objective
The total pressure of a mixture of gases is equal to the sum of the individual partial pressure of the gases.

10.4.1.1 Essential Knowledge
Dalton's Law of Partial Pressures says that the total pressure of a mixture of gases will be equal to the sum of the partial pressures of the gases that make up that mixture. This is often used when collecting a gas over water. The total pressure of the gas collected over water is equal to the pressure of the gas in addition to the pressure of the water vapor.

OPTIONAL

Topic 10.5 Gas Laws and Chemical Reactions

10.5.1 Learning Objective
Stoichiometry can be used to make predictions of amounts formed in a normal chemical reaction that involves one or more gases.

10.5.1.1 Essential Knowledge
The Combined Gas Law and Ideal Gas Law can both be used with stoichiometry calculations.

Unit 11. Solutions

Topic 11.1 Solutions

11.1.1 Learning Objective
IMF can be used to determine if a solute and solvent will form a solution.

11.1.1.1 Essential Knowledge
A solution is a homogeneous mixture made up of a solute and solvent. The solute is the substance being dissolved and is present in a lesser amount while the solvent is the substance doing the dissolving and is present in a greater amount. The solute and solvent can be solids, liquids, or gases. The IMF of the solute and solvent determines if a solution will form. If the solute and solvent have similar IMF a solution will form. If the solute and solvent have dissimilar IMF a solution will not form. In general, a polar solute will dissolve in a polar solvent and a nonpolar solute will dissolve in a nonpolar solvent, but polar and nonpolar solute/solvent will not be soluble.

Topic 11.2 Electrolytes and Nonelectrolytes

11.2.1 Learning Objective
A solute can be an electrolyte or nonelectrolyte.

11.2.1.1 Essential Knowledge
An electrolyte is a solute that will conduct electricity and light up a bulb when dissolved in water while a nonelectrolyte will not conduct electricity when dissolved in water. Ionic compounds are electrolytes and polar covalent compounds are nonelectrolytes.

11.2.1.2 Exclusion Statement
The difference between strong and weak electrolytes will not be assessed.

Topic 11.3 Solubility

11.3.1 Learning Objective
A solubility chart gives the solubility of different compounds in 100 g of water at varying temperatures.

11.3.1.1 Essential Knowledge
A saturated solution has as much solute dissolved as possible at that temperature and volume of water. A solubility chart shows the mass of a substance required to make a saturated solution in 100 g of water at varying temperatures. The line of each substance is the point of saturation; any mass of substance greater than the line is considered saturated and any mass of substance less than the line is considered unsaturated at that temperature. When the temperature changes the solubility, and hence the point of saturation, also changes.

11.3.1.2 Exclusion Statement
Supersaturated solutions will not be assessed. The mass of solid required for saturation for amounts of water less than or greater than 100 g will not be assessed.

Topic 11.4 Concentrations of Solutions

11.4.1 Learning Objective
Molarity is used in science to discuss the concentration of a solution.

11.4.1.1 Essential Knowledge
The concentration of a solution can be generally discussed as being concentrated or dilute. However, molarity is used in science to specifically discuss concentration.

$$Molarity = \frac{moles\ of\ solute}{Total\ Volume\ of\ Solution}$$

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

where:

M = molarity (mols/L)
n = moles of solute (moles)
V = total volume of solution (L)

Topic 11.5 Dilution and Chemical Reactions

11.5.1 Learning Objective

A solution can be diluted thus changing the molarity of the solution.

11.5.1.1 Essential Knowledge

A higher molarity solution can be diluted and made into a lower molarity solution by adding distilled water. When doing so, the moles of the higher molarity and lower molarity solution remains the same. This allows the use of the equation given below to determine the required molarity of a stock solution or the diluted molarity of a desired solution.

$$M_1V_1 = M_2V_2$$

where:
M_1 = Initial higher molarity (stock solution)
V_1 = Volume of M_1
M_2 = Desired lower molarity
V_2 = Volume of M_2

Unit 12. Reaction Rates

Topic 12.1 Rates of Reaction

12.1.1 Learning Objective

The collision theory can be used to explain the rate of a reaction.

12.1.1.1 Essential Knowledge

There are three postulates to the collision theory:

1. Molecules must collide
2. Molecules must have the proper orientation upon colliding
3. The molecules must have a minimum amount of energy, termed the activation energy

The reaction rate, the amount of time for the reaction to occur, can be changed by changing the concentration, temperature, or surface area of a reactant. A catalyst can also be used to increase the rate of a reaction.

Unit 13. Acids and Bases

Topic 13.1 Acids, Bases, and Their Definition

13.1.1 Learning Objective

Acids and bases have distinct properties.

13.1.1.1 Essential Knowledge

Acids have a sour taste, change blue litmus paper red, are colorless in the presence of phenolphthalein, and corrode some metals. Bases have a bitter taste, change red litmus blue, and are pink in the presence of phenolphthalein.

13.1.2 Learning Objective

Determine the acid, base, conjugate acid, and conjugate base in a chemical reaction.

13.1.2.1 Essential Knowledge

Acids donate a hydrogen ion, H^+, and bases accept a hydrogen ion. The deprotonate acid is the conjugate base and the protonated base is the conjugate acid. It can also be said that the conjugate base accepts the hydrogen ion and the conjugate acid donates the hydrogen ion when moving from products to reactants.

13.1.2.2 Exclusion Statement

The difference between Bonsted-Lowry and Arrhenius acids and bases will not be assessed. Lewis acids and bases will not be assessed.

Topic 13.2 The pH Scale

13.2.1 Learning Objective

Calculate the pH and pOH of a strong acid and strong base solution in order to be able to determine if the solution is an acid, base, or neutral.

13.2.1.1 Essential Knowledge

The concentration of hydrogen (or hydronium) ion and hydroxide ion are often reported as pH and pOH respectively.

$$pH = -\log [H^+]$$

$$pOH = -\log [OH^-]$$

Acids have a pH below 7.00 and bases have a pH above 7.00. A neutral solution has a pH of 7.00. The pH and pOH are related to each other according to the equation below:

$$pH + pOH = 14$$

13.2.1.2 Exclusion Statement

Weak acids and bases will not be assessed.

Topic 13.3 Acid-Base Titrations

13.3.1 Learning Objective

The molarity of a monoprotic acid can be determined via titration.

13.3.1.1 Essential Knowledge

An acid-base reaction can be carried out under controlled conditions in a titration. At the equivalence point, the number of moles of titrant added is equal to the number of moles of analyte originally present. This relationship can be used to obtain the concentration of the analyte.

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