

Chemistry

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Unit I: Atomic Structure and the Periodic Table

Lesson 1: The Historical Development of Atomic Theory

Topic 1.1: Dalton's Atomic Theory

Democritus (Greek natural philosopher) was first to hypothesize the existence of an atom – non-scientifically.

In 1800s, John Dalton & Joseph Proust used scientific method to research the atom, developing the first atomic theory (an explanation of the structure of matter in terms of different combinations of very small particles)

Proust discovered that compounds follow the “law of definite proportions:” whole numbers define the ratio of masses of elements found in a compound

Law of definite proportions allowed Dalton to form atomic theory (which supported law of definite proportions)

Proust split tin oxides to determine ratios, but Dalton combined tin + oxygen to measure.

Tin oxide forms in two different types, and Dalton observed this by testing the amount of oxygen with which tin reacts.

Dalton's Atomic Theory:

- Atoms are indivisible; atoms are the smallest components of matter and cannot be broken down.
- Elements are composed of identical atoms and these atoms are unique to each element; atoms of different elements are different.
- Compounds are composed of atoms of two or more elements in fixed proportions; the ratio of atoms in a compound is fixed for all samples of the compound.
- A chemical reaction is a rearrangement of the atoms into new combinations; the reaction cannot create or destroy atoms, only rearrange

Note: Dalton's Atomic Theory was incomplete like nearly every other theory.

Topic 1.2: Discovery of the Electron

J.J. Thomson discovered electron in late 1800s w/ cathode ray tube experiments (electron emitters)

A stream of particles (electrons) was emitted through a tube, and when it reached the other end deflected towards a positive electric charge.

Electrons were clearly smaller than atoms, so Thomson concluded that they must be part of an atom and atoms have subatomic particles (disproving Dalton's indivisibility of atoms assertion)

Topic 1.3: Plum Pudding Model

Thomson's results caused the proposal of the plum pudding model, which defined a model such that

- Electrons float in a sea of positive charge
- Only slight modification
- Recognized existence of electrons, but assumed that positive charge counterbalances them, and the atom is thus neutral

Robert Millikan conducted an oil drop experiment where droplets were charged w/ spray and a single droplet was held in mid-air. Allowing the oil to drop, the force of gravity was compared to the acceleration of the droplet and the electric field, and it was determined (because of the quantization of charges on the oil droplets to multiples of the charge of an electron) that electrons have $1.6 * 10^{-19}C$ of charge. Applying this to Thomson's ratio of mass-to-charge for an electron, $\frac{1g}{1.76*10^8C}$, a mass of $9.09 * 10^{-28}g$ was obtained.

Topic 1.4: Discovery of the Nucleus

In a test of the plum pudding model by Ernest Rutherford, positively charged particles shoot out of a particle gun towards a thin sheet of gold foil. Then, reflected particles or particles passing through the foil are detected by a surrounding screen.

The results of Rutherford's experiment:

- Most particles passed straight through.
- A few particles are deflected at large angles

He concluded that an atom is mostly empty space, with a small, positively charged, central region (the nucleus), in which most of the atom's mass is concentrated.

Lesson 2: The Modern Atomic Theory

Topic 2.1: Theories of Light

Newton, in 1704, developed the corpuscular (from Latin for "puny body") theory of light, light as particles.

Thomas Young in the early 1800s, used the double-slit diffraction experiment to demonstrate light bending around corners and interfering like a wave.

In 1887, Hertz showed that light shining on metal produces emissions (photoelectric effect) by knocking electrons off of the metal – contradicting the wave theory of light. Thomson showed that those emitted particles are electrons. He also discovered that the energy of electrons depends on frequency of the light, not intensity.

Topic 2.2: Einstein and the Photoelectric Effect

Einstein noticed that electrons came off in quantized packets (quantum). Borrowing Max Planck's idea that vibrating molecules only had energy at certain values, he treated light particles as quantum where energy is given by $h\nu$ where h is Planck's constant and ν is the frequency of light. This explained the correlation of light frequency with electron energy but not quantity and light intensity's correlation with electron quantity but not energy. This was the start of Quantum Theory.

Topic 2.3: Discrete Emission Lines

Johannes Rydberg studied the emission spectra of heated metals.

An emission spectrum is a visible light in which wavelengths of light emitted by a substance while heated show up as bright, colored lines.

Surprisingly, the spectra were discrete instead of continuous. This allowed Rydberg to establish equations relating frequency and energy of light.

Building off of this, Niels Bohr (1915) defined a new model for the atom (the planetary model) with electrons orbiting the nucleus in fixed, discrete energy-level paths (i.e. electrons can only exist in specific orbits and in no other location). This model was a modified version of Rutherford's model of the atom (a central nucleus with electrons positioned in any position around the nucleus – not discrete energy levels)

In the planetary model, each of the orbits has a specific energy amount and higher orbits have higher energies. Also, according to this model, electrons can't have internal energy (completely determined by its orbit). Electrons jump between energy levels by absorbing or emitting light, but because there are a discrete number of orbits, the frequencies of the light are discrete (thus defining emission lines).

Topic 2.4: Problems with the Bohr Model and the Electron Cloud Model

The Bohr Model explains a lot of data, but it still has a number of problems:

- It doesn't explain how some spectral lines are brighter than others, and they can come in groups of multiples
- The model doesn't work for Helium or any larger atoms' spectral lines
- Electrons have a definite radius and momentum, contradicting the Heisenberg Uncertainty Principle.

The electron cloud model (developed by Schrodinger and others) replaced the Bohr Model, and is the currently accepted model, because it considered the Heisenberg Uncertainty Principle. In this model, electrons are considered to only have probable locations within a "cloud," but the clouds still have different energy levels for the electrons.

Lesson 3: Atomic Spectra

Topic 3.1: Bohr's Explanation of Atomic Emission and Absorption

Electrons aren't strictly in orbits; instead, they rest in discrete *orbitals*, or energy levels. Electrons absorb light ("excitation" of an atom) and rise to higher energy levels and randomly fall to lower energy levels, emitting light. However, the Bohr model can represent this sufficiently.

This also explains the quantization of atomic spectrum lines (there are a finite number of "jumps" electrons can make between orbitals)

Topic 3.2: Absorption and Emission Spectra

Atomic spectra are the frequencies of light at which atoms' electrons "jump" between orbitals. These are visible in "absorption" and "emission" spectra.

Absorption spectra is an EM spectrum in which wavelengths of light absorbed by an element show up as dark lines on a visible light background. Emission spectrum is the complement – brightly colored lines on a black background. These are determined to represent all the possible emission/absorption frequencies of a given atom, respectively.

Topic 3.3: Spectroscopy

Spectroscopy (the study of absorption and emission of light by atoms) applies these principles and observations by:

- identifying the composition of the sun and stars
- identifying elements in a sample
- determining chemical bonds
- identifying harmful bacteria in food and environmental samples
- studying surfaces of materials

Lesson 4: The Structure of the Atom

Topic 4.1: The Atom's Components

An atom is the smallest particle of an element that has the same properties as the element. The atom can be divided into two parts, the nucleus (the center that holds protons and neutrons), and the orbitals (regions surrounding the nucleus where electrons sit in "clouds")

The atom is made of 3 particles:

- Protons: +1 charge, mass of 1 proton, in nucleus
- Neutrons: 0 charge, mass of 1 proton, in nucleus
- Electron: -1 charge, mass of 1/1800 proton, in electron cloud/orbitals

In a neutral atom, # of protons = # of electrons, but neutrons can vary significantly. Any atom can have any number of these particles, but it's not always stable.

Topic 4.2: Atomic Mass Unit

Because atoms' masses are so small, scientists use a unit called an "AMU" to describe weights of subatomic particles.

An AMU is defined to be exactly $\frac{1}{12}$ of the mass of a C_{12} atom, which is equal to $1.66 \times 10^{-24}g$.

The mass of a proton is $1.673 \times 10^{-24}g \sim 1amu$. The mass of a neutron is $1.675 \times 10^{-24}g \sim 1amu$. The mass of an electron, however, is $9.109 \times 10^{-28}g \sim 0.0006amu$.

Notably, protons and neutrons are not the most basic subatomic particle. They are, in turn, composed of quarks.

Topic 4.3: Atomic Number

As all atoms can contain any number of subatomic particles (neutrons, protons, and electrons).

However, not all elements can contain any number of subatomic particles. In fact, all atoms of a given element are restricted to their atomic number (Z) of protons. For example, aluminum has an atomic number of 13, therefore it always has 13 protons, by definition. Also, because this property is unique to an element, no two elements share the same atomic number/number of protons. On the periodic table, these increase along a row.

Topic 4.4: Ions

The number of electrons, unlike protons, can vary between atoms of the same element. These amounts are often changed in chemical reactions. For example, aluminum could gain or lose several electrons and still remain aluminum. This would make it an ion, and dependent on the number gained or lost, it would become more negatively or positively charged (respectively because electrons have a -1 charge).

Topic 4.5: Isotopes

Mass Number (A) is the number of protons and neutrons in a given atom. This can vary from atom to atom of the same element. Atoms of the same element and same mass number are isotopes.

There are several ways to represent it (all of these are an aluminum atom with a mass number of 27): aluminum-27, Al-27, or ^{27}Al .

If both Mass Number and Atomic Number must be displayed, the mass can be displayed with a superscript and the atomic w/ a subscript (Figure 1).

Where atomic number (Z) and mass number (A) are known, the number of neutrons (N) follows $N = A - Z$



Figure 1: Isotope Representation

The mass number on the periodic table is a weighted average (from abundance of isotopes) of available isotopes in nature. For example, Cl-35 (35amu) shows up in nature 75.78% of the time and Cl-37 (37amu) shows up 24.22% of the time. The average is therefore $35(75.78\%) + 37(24.22) \simeq 35.45\text{amu}$

Lesson 5: Nuclear Fission and Nuclear Fusion

Topic 5.1: Nuclear Decay

There are two main types of radioactive decay:

- Alpha Decay: A parent nucleus emits an He4 (alpha) particle, losing 2 protons and 2 neutrons. -2 atomic number and -4 mass.
- Beta decay
 - Beta-plus decay: A parent nucleus proton becomes a neutron, emitting a positron. Mass stays the same, and -1 atomic number
 - Beta-minus decay: A parent nucleus neutron becomes a proton, emitting an electron. Mass stays the same, and +1 atomic number

In all nuclear (fission, fusion, and decay) reactions, the total mass and total atomic number (protons) stay the same. This allows for algebraic completion of any nuclear reaction. Note that this includes subatomic particles, with the respective nuclear formulae:

- Electron: ${}_{-1}^0e$
- Neutron: ${}_0^1n$
- Proton: ${}_1^1p$

Topic 5.2: Fission

“Nuclear fission is the process in which a heavy nucleus is split into two large fragments of comparable mass to form more stable and smaller nuclei, resulting in the release of great amounts of energy.” Nuclear fission works by a neutron hitting a large nucleus and making it into an unstable isotope. Then, the resulting isotope quickly decays to fission products, a large amount of energy, and neutrons.

This isn't a completely clean energy source, however. Nuclear energy processes' large fission products are radioactive waste which needs to be stored far away from humans. These can occur naturally, but they don't occur explosively. In man-made power plants or nuclear bombs, fission reactions self-sustain

because the output neutrons from one fission reaction trigger more fission reactions. These reactions only self-sustain if the mass of the reactive material is at or above critical mass, which is different for different nuclei.

Example fission pathway: ${}_{92}^{235}\text{U} + {}_0^1\text{n} \rightarrow {}_{38}^{94}\text{Sr} + {}_{54}^{140}\text{Xe} + 2{}_0^1\text{n}$

Topic 5.3: Fusion

"Nuclear fusion is the process in which lighter atomic nuclei combine to form a more stable heavier nucleus, resulting in the release of great amounts of energy." If humans were able to harness the power of fusion reactions, then these would be very useful because they produce no nuclear waste. However, we can't because they require very high activation energies, and containment materials can't withstand the reaction long-term. This is why the only known location of fusion reactions is within stars.

A common fusion pathway in stars (including our sun) is ${}^2\text{H} + {}^3\text{H} \rightarrow {}^5\text{He} + 17.6\text{MeV} + {}^4\text{He} + \text{n}$. Another common pathway is the Carbon-Oxygen-Nitrogen cycle, which uses those atoms to turn 4 independent protons into a ${}^4\text{He}$ nucleus and a significant amount of energy.

Topic 5.4: Cold Fusion

This is a hypothetical fusion mechanism which would have lower activation energies than a standard "hot" fusion reaction (roughly room temperature). The hypothetical pathway involves deuterium (H-2) water with palladium and platinum. It would produce minimal waste.

Sadly, this is controversial and most scientists think that cold fusion isn't possible. This is controversial because there is one team claiming the ability to perform cold fusion, but they neglected to measure helium, their neutron production results were inconsistent with expected numbers, and the results of a similar experiment with conflicting and inconclusive results was completely ignored. Foremost, they didn't allow the scientific community to peer review their results.

Lesson 6: Elements, Compounds, and Mixtures

Topic 6.1: Physical vs Chemical Properties/Changes

Physical properties of a substance can be examined without changing the chemical structure of the substance. Chemical properties, on the other hand, cannot be. This is the sole and singular difference. Physical/chemical changes are the respective changes of such properties.

Topic 6.2: Classification of Matter

All matter is considered either a pure substance (any matter that cannot be decomposed into simpler components without a chemical change) such as water

or salt or a mixture (a combination of two or more pure substances that are not chemically combined) such as salt water, tea, steel, or air.

Elements are a pure substance of one type of atom such as Hydrogen, Helium, Oxygen, Carbon, Iron, or Chlorine. Oxygen's most stable forms are molecular (O_2 and O_3), but this is still considered an elemental form.

Compounds are another pure substance – solely composed of one molecule – such as table salt ($NaCl$), ammonia (NH_3), water, emerald, or rust.

Mixtures are also divisible into two types: homogeneous and heterogeneous mixtures. Homogeneous mixtures are indistinguishable between any two regions and appear as one phase. Heterogeneous mixtures appear as two or more phases, and are different between at least two regions.

Topic 6.3: Separation of Mixtures

Separations of mixtures use physical methods only (because, by definition, mixtures can be separated without chemical methods). These include:

- Sorting: manual separation of such a mixture as discrete physical objects
- Filtration: separates a mixture by pushing through one part but leaving behind the other (such as suspended solids)
- Distillation: a difference in boiling points allows one part to be boiled away, such as boiling gasoline out of crude oil.
- Chromatography: Passing a mixture in solution, suspension, or vapor form through a medium in which the components move at different rates. Named after dyes "climbing" through hanging solvent-soaked paper at different rates and thus separating

These all take advantage of the differences of physical properties of materials such as solubility, boiling point, or size.

Lesson 7: Atomic Numbers and Electron Configurations

Topic 7.1: Quantum Numbers

In the electron cloud model, electrons are located in orbitals (a given location where an electron can be). A quantum number describes the location of the electron. n (principal quantum number) is the orbital size. l (angular momentum quantum number) is the orbital shape. m (magnetic quantum number) is the orbital orientation. Quantum numbers are constrained to these values:

Quantum No.	Possible Values
n	Nonzero positive integers
l	Positive integers between 0 and $n-1$
m	Integers between $-l$ and $+l$

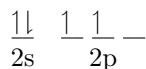


Figure 2: Lithium's Orbital Notation

Topic 7.2: Shells, Subshells, and Orbitals

A shell is a set of orbitals with the same principal number, n . In a non-excited state – as is the default for atoms – electrons occupy the lowest energy orbitals and fill randomly between orbitals of equal energy. This manifests in shells filling consecutively. Full shells are more stable (I would recheck this, and it's likely not completely true). Bohr diagrams are often used to represent these shells.

A subshell is all the orbitals with the same principal (n) and angular momentum (l) quantum numbers. Similarly, an orbital has unique set of principal, angular momentum, and magnetic quantum numbers.

Shells are represented by numbers (i.e. shell with a principal number n is shell n). Subshells are represented by a number and a letter. The angular momentum number l follows this pattern (e.g. a subshell where $n=3$, $l=0$ is $3s$):

l	letter
0	s
1	p
2	d
3	f

Topic 7.3: The Pauli Exclusion Principle

Quantum mechanics dictates that no two fermions (for our cases, electrons) can share the same quantum numbers (the only defining characteristics of a given fermion), so any given orbital, subshell, or shell can hold a finite number of electrons.

For an orbital, the maximum number of electrons is always 2 because of an as of yet unmentioned quantum number, spin. There are two possible states of spin ($-\frac{1}{2}$ and $\frac{1}{2}$), which allow the electrons to be unique. Going up a level to subshells, one can tell that there are $2l + 1$ orbitals ($4l + 2$ maximum electrons), a direct result of the constraint of the magnetic quantum number, $-l \leq m \leq l$. Going up another layer, a shell contains n subshells, but these have an unequal number of orbitals. It contains n^2 , which can be determined after short mathematical manipulation (and consequently $2n^2$ electrons)

Topic 7.4: Electron Configuration and Orbital Notation

Orbital notation (see figure 2) is a type of notation which shows how electrons are positioned in an atom. Lines represent orbitals, numbers and letters on the bottom represent subshells, and arrows represent electrons and their spin.

Electron Configuration is another type of notation which is more concise but still communicates which subshells are filled. It reads like $1s^1$ or $1s^22s^2$

Hund's Rule determines the order in which orbitals of the same subshell fill. All orbitals fill in order, as is convention, singly (one electron in each) in order to maximize atomic stability and separate electrons from one another. After all orbitals are filled, orbitals start to fill doubly. Note that all singly filled orbitals' electrons have the same spin.

Topic 7.5: Aufbau Principle

The Aufbau Principle states that electron subshells fill in order of energy, instead of numerically. This exception shows up with the d and f orbitals. This is because, for example, the $E_{3d} > E_{4s}$.

The **diagonal rule** identifies these exceptions (and the general rule) of how subshells energies equal. Following through the arrows in order gives increasing energy of subshells.

Topic 7.6: Dot Structures/Bohr Diagrams

Dot Structures look like the simplified Bohr model of the atom. There is a single, monolithic nucleus in the middle (labeled with the element), and rings or circles spaced around the atom (representing shells). For every ring, there are a number of electrons which appear on it as dots. These are useful for readily representing the high-level structure of an atom.

Lesson 8: The History and Arrangement of the Periodic Table

Topic 8.1: The Periodic Table

The periodic table is an organized display of the elements. On the modern periodic table, elements are arranged in order of increasing atomic number. They are grouped by similar chemical properties and electrical configurations.

Antoine Lavoisier wrote the *Elementary Treatise of Chemistry* in 1789, the first modern chemical text book. He classified substances/elements in 4 groups: acid-making (sulphur, phosphorous), gas-like (light, heat, oxygen, nitrogen), metallic (cobalt, mercury, tin, gold, silver, manganese), and earthy (calcium oxide, magnesium oxide, silicon dioxide) elements. Note that several of these are incorrectly identified as elements (they're either not matter or compounds)

Topic 8.2: Early Attempts at a Periodic Table

John Dobereiner (1829) designed an early periodic table with "triads." He arranged elements with similar properties into groups of 3. He chose elements with similar mass gaps (i.e. the smallest and middle elements had a similar difference in mass as the middle and largest elements).

Building off of this, another scientist, John Newlands, developed the law of octaves. Among known elements, he discovered that a given element has similar properties to an element eight spaces separated if the elements were ordered by atomic mass.

Dmitri Mendeleev (1869) designed a complete periodic table of all known elements. He ordered elements by increasing atomic mass, and included blank spaces for undiscovered elements based on similarities/patterns of chemical properties.

Topic 8.3: The Modern Periodic Table

Henry Moseley (1909) completed the final revision to today's periodic table by using atomic number, instead of atomic mass, to organize the elements because this accounted for mass variations of different isotopes.

Topic 8.4: Metals, Nonmetals, and Semimetals

On the periodic table, there are three major groups of elements: metals, nonmetals, and semimetals or metalloids.

Metals tend to be malleable (easily pounded into a sheet), ductile (easily drawn into a wire), conductive, and solid. Nonmetals tend to be the opposite – brittle, insulating, and solid, liquid, or gas. Semimetals have the properties of both.

Topic 8.5: Periods and Groups

Periods are the rows of the table (because they occur with the same properties in order, repetitively). The atomic number increases from left to right, and chemical properties change systematically across the table.

Groups or families are groups of chemical properties, i.e. they all have similar chemical properties. These appear as columns and are labeled 1-18 or 1A-8B.

On the periodic table, information about the individual elements is also easily visually available: the name, its symbol, its atomic number and its mass number.

Topic 8.6: Minor Groups

Groups on the periodic table are often have significant chemical properties and are thus named.

The alkali metals in group 1 or 1A are typically:

- Silver in color
- Soft (can be cut with a knife)
- Highly reactive with oxygen and water
- Able to oxidize in air

- +1 cation because they can easily lose an electron

The alkaline earth metals in group 2 or 2A are typically:

- Silver in color
- More brittle than alkali metals
- Somewhat reactive
- Low in density, with low melting and boiling points
- +2 cation because they lose 2 electrons

The halogens in group 17 or 7A are nonmetals which are typically:

- Highly reactive with metals
- Toxic to biological organisms
- Naturally occurring as diatomic molecules (F_2 , Cl_2 , Br_2 , I_2)
- Highly reactive with alkali and alkaline earth metals to form salt
- -1 anion because they gain 1 electron

The noble gases in group 18 or 8A are nonmetals which typically:

- Are completely inert gases (nonreactive)
- Are odorless and tasteless
- Are nonflammable
- Have an extremely low boiling point
- Produce characteristic colors when excited electrically

The transition metals in groups 3 to 10 or B typically form colored compounds, and can have useful properties like magnetism and high conductivity.

The inner transition metals (actinides and lanthanides) are a part of the transition metals which appear at the bottom of the table. These are often radioactive, especially elements after Uranium. Any element after Uranium is called "trans-Uranium," and these elements are only synthesised with nuclear accelerators, other than trace amounts on Earth

Lesson 9: Electrons and the Periodic Table

Topic 9.1: Noble-Gas Notation

Every element has an electron configuration, including noble gases. Because noble gases have full outer shells (e.g. Helium has $1s^2$ configuration, Neon has $1s^2 2s^2 2p^6$), electron configurations can be shortened by referencing the "closest" noble gas. For example, Sodium's electron configuration is $1s^2 2s^2 2p^6 3s^1 = [Ne]3s^1$. In contrast, "normal" notation is called longhand.

Topic 9.2: Core and Valence Electrons

Noble-Gas Notation is also useful because it emphasizes valence electrons, the chemically most reactive part. The electrons which are part of the abbreviation (such as $[Ne]$) are called core electrons and are generally unreactive because they are part of a full shell (much like the noble gases they are abbreviated to).

A notable exception to the rule of everything after the gas being a valence electron is the "out of order" subshells. A full subshell part of a lower shell than the other subshells (e.g. $3d^{10}$ in $4s^23d^{10}4p^3$) is part of the core instead of the valence. This does not apply to partially filled "out of order" subshells.

Topic 9.3: Valence Electron Blocks

When looking at elements in sequence, valence electrons' orbitals fill up until there is a subshell which is only partially filled. This highest energy subshell determines the "block" of the element. For example, an element with an electron configuration $1s^22s^22p^4$ is in the p-block. These form "blocks" across the table, which can allow faster determination of noble-gas configurations (when doing this, take note of "out of order" subshells)

Some elements do not follow the "correct" diagonal rule trends, mostly in the d-block, f-block, and transitional metals. However, knowing the exceptions is unnecessary.

Lesson 10: Periodic Trends

Topic 10.1: Trends of Atomic Radii

Atomic radius is "half the distance between two identical atoms in a diatomic molecule." Because electrons in an energy level "shield" the nuclear attraction from others, atomic radii tend to decrease across a period. Atomic radii also tend to significantly increase down a group. This means that the largest atomic radius is in the bottom left corner.

Topic 10.2: Trends of Ionic Radii

Ionic radius is "a measure of the size of an ion." An atom has a "standard" ion (such as Al^{3+} for Aluminum) which is generated when it tries to form a full valence shell. The ionic radius compared to the atomic radius is larger for anions (negative) but smaller for cations (positive). Similar to atomic radii, ionic radii increase down a group and decrease along a period within the cations or anions. Significantly, there is a slight increase between cations and anions (e.g. $r_{B^{3+}} < r_{C^{4-}}$)

Topic 10.3: Transition Metal Exceptions

Ionic radii and atomic radii "rules" are not unbreakable. Transition metals often break these rules because they fill the d subshells, which is in a smaller

radius shell. This makes atomic radii increase across a period after a slight decrease. The cation radii also vary, but the pattern is irregular.

Topic 10.4: Ionization Energy

To form the ions mentioned in topic Trends of Ionic Radii, there is a significant amount of energy required to remove the electrons. This amount of energy (for the gaseous form of an element) is called the ionization energy. The ionization energy depends on several factors: the nuclear charge (a stronger nucleus is harder to remove from), the distance of the electron from the nucleus (a far away electron takes more energy), and the number of electrons removed (removing an electron while there is a strong electric charge is more difficult than removing from a neutral atom). Because of the noticeable difference in the number of electrons removed, first, second, third, etc. ionization energies are recorded as the energy required to remove a given electron. The first ionization energy tends to:

- increase across a period (nuclear charge increases)
- slightly decrease between groups 2 and 13 (begins to add to new subshell) & groups 15 and 16 (begins to add to half-full subshell)
- decrease across each group (higher nuclear distance)

See graph at [Lumen Learning \(fig. 1\)](#)

Topic 10.5: Trends in Electron Affinity

Electron affinity is the energy required to add an electron to a neutral atom in the gas phase. Often, this amount is negative because adding an electron releases energy. Electron affinity generally becomes more negative (more energy released) across a period, but there are several exceptions. For example, groups 2 and 15 release very small amounts of energy because a new or half-full subshell begins to be filled. It also tends to become less negative down a given group because the nuclear distance increases, and the potential energy difference becomes smaller.

Topic 10.6: Trends in Electronegativity

When atoms bond, they pull on valence electrons shared in the bond. Electronegativity measures the strength of this pull. This tends to increase across a period (higher nuclear charge) and decrease across a group (higher nuclear distance)

Unit II: States and Properties of Matter

Lesson 1: Gases

Topic 1.1: Kinetic-Molecular Theory

Kinetic-molecular theory is a model of gases “as a large number of constantly and randomly moving particles that collide with each other and the walls of the container.”

The model postulates that:

- Gases are made of a large number of hard spheres (like billiard balls) in continuous, random motion
- Most of the volume of the gas is empty
- There is no force of attraction or repulsion between gas particles
- All collisions between particles are perfectly elastic (no energy is lost)
- Average kinetic energy is only dependent on temperature.

An ideal gas is the theoretical gas that is well-modeled by kinetic-molecular theory. Under common conditions, most gases are approximately ideal.

This also allows the relation of pressure (force per unit area of the container), volume (size of the container), temperature (average kinetic energy), and number of molecules (amount of gas) for the same gas. $\frac{P_1 V_1}{N_1 T_1} = \frac{P_2 V_2}{N_2 T_2}$.

Topic 1.2: Properties of Gases

Most gases are compressible (a gas’s container can shrink) and expand to fill their container. This is well explained with kinetic-molecular theory – with large spaces between particles and no interactions between them, a smaller container would make any given particle stay within a smaller space (the gas is already mostly space), and the lack of interaction allow them to spread out as they will.

In a gas, particles move randomly, so if there is no barrier between a high concentration and low concentration region, diffusion will occur, making the entire volume a medium concentration region.

Similarly, if there is a very small opening between high and low concentration regions, the particles will effuse through the hole, bringing the concentration to an equilibrium. Notably, the rate at which this occurs is determinable with Graham’s Law: $r \propto 1/\sqrt{d}$. This is a natural consequence of the assumption by kinetic-molecular theory that average kinetic energy of the gas is only dependent upon temperature (and is thus constant between two gases of different densities with all other properties held the same)

Lesson 2: Liquids

Topic 2.1: Particles in Liquids

In liquids, kinetic-molecular theory is still obeyed, but one of the gas postulates is different: liquids' intermolecular attractions are not negligible. Instead, they are comparable with intermolecular attractions. As a consequence, this means they are denser than gases and have a fixed volume (incompressible). However, the particles aren't fixed in place by intermolecular forces, so they are able to flow and take the shape of the container.

Topic 2.2: Intermolecular Forces and State Changes

Intermolecular forces affect interactions between particles. The strength of these forces changes condensation/evaporation points. This is because if $KE > F$, the material is gaseous, but if $KE < F$, the material is liquid, so as F increases, the KE threshold/evaporation point increases.

A similar pattern can be observed with freezing/melting. Higher intermolecular forces causes a higher the freezing/melting point to increase.

Topic 2.3: Viscosity

Viscosity is the thickness or resistance to flow of a liquid. This occurs because of significant intermolecular forces. On this same line of thinking, higher kinetic energy (temperature) is inversely correlated with viscosity. For comparisons between liquids, smaller (less massive) particles can move easier, and liquids with weaker intermolecular forces can also move more easily.

Topic 2.4: Surface Tension

Surface tension is the tendency of a liquid to resist penetration. This is stronger than standard viscosity because it is a 2 dimensional force instead of a 3 dimensional force. Like viscosity, this is correlated (directly) with intermolecular attraction. A surfactant such as dish soap can reduce surface tension by breaking up the "skin" of molecules.

Topic 2.5: Other Properties

Liquids are generally incompressible. This allows them to transmit forces such as in hydraulics.

Liquids can also dissolve solids, other liquids, and gases. When dissolving another material, particles become evenly dispersed among liquid particles. When two liquids dissolve, they are "miscible" (able to dissolve into eachother). This is in contrast to immiscible liquids like oil and water which remain separate. As mentioned earlier, surfactants can "connect" immiscible liquids like water and oil.

Lesson 3: Solids and Plasmas

Topic 3.1: Crystals and Amorphous Solids

Solids are a cold state of matter which are rigid and strongly resist change to shape or volume. Some solids are crystalline, meaning that they have a regular (repeating unit which is significantly larger than the particles themselves) lattice arrangement. Crystalline solids are incompressible (typically dense and hard) and vibrate.

Amorphous solids, on the other hand, are not periodic and crystalline, but these are still solid. The particles often form amorphous structures from rapid cooling which halts assembly of unit cells as in crystals. This can allow liquid-like movement

Topic 3.2: Plasma Properties

A plasma is the state of matter hotter than a gas. The difference is that plasmas have ionized particles (i.e. electrons are detached from atoms by energy). The temperature of thermal plasmas (the most common variety) can be thousands of degrees or more. They conduct electricity, are compressible, and have no definite shape or volume like gases.

Topic 3.3: Plasma in Nature and in Technology

When charges build up from clouds to ground, it's so powerful that plasma forms along its path in order to discharge the electrons through a lightning strike. Both the sun and the stars are plasmas, and the solar wind's interactions with the magnetosphere and upper atmosphere forming auroras are plasmas.

Technology utilising plasmas are very common: fluorescent lights, ion thrusters, arc welders, plasma displays, and novelty plasma balls. Plasma displays and plasma balls pose an interesting exception to "standard" thermal plasmas (such as lightning, stars, or arc welders) – cold (non-thermal) plasmas. These are able to exist as cold plasmas because:

- They often have hot/fast electrons (which are exceedingly light) but cold (slow) nuclei allowing a thermal equilibrium.
- Electrons' kinetic energy becomes light
- Fewer than 1% of particles are ionized

Lesson 4: Phase Changes

Topic 4.1: Molecular Motion and Phase Changes

Phases of matter follow a strict hierarchy of energy: $E_{solid} < E_{liquid} < E_{gas} < E_{plasma}$. Between the coldest 3 state of energy, any given transition is possible. These are named as Melting/Freezing between solid and liquid, Boiling/Condensing between liquid and gas, and Sublimation/Deposition between

solid and gas. The first of the phase changes in each pair occur when a substance absorbs energy. Conversely, the latter phase change in these pairs makes substances release energy.

There are some notable features of boiling/condensing within a closed container, related to the equilibrium between liquid and gas. This occurs because particles are constantly moving between liquid and gas phase, and the gas exerts a pressure (vapor pressure) on the liquid's surface, which changes the liquid's evaporation point.

Topic 4.2: Diagramming Phase Changes

As energy is added to a substance, temperature does not increase linearly. In fact, during phase transitions, temperature remains constant at the melting/boiling points. When not in a phase transition, temperature increases directly with energy.

Lesson 5: Changes in Matter

Topic 5.1: Matter

Chemistry studies matter – anything that has mass and takes up space. Matter can take any phase of matter, and is virtually everywhere. In a kitchen, for example, the chef, ingredients, utensils, and air are matter.

Topic 5.2: Physical and Chemical Properties

Physical properties are characteristics of a substance which can be observed without changing the identity (chemical structure) of the substance. These are numerous, including: color, odor, texture, melting/boiling point, state of matter, mass, weight, density, transparency, malleability, magnetism, and conductivity.

In contrast, chemical properties are characteristics only observable when the substance interacts with another substance. These include flammability and reactivity.

To determine if a given property is chemical or physical, one can ask a couple of basic questions about a given property: "Must multiple substances be present?" (if yes, chemical) and "Does a new substance form?" (if yes, chemical)

Topic 5.3: Extensive and Intensive Physical Properties

Physical properties are further divided into two categories: intensive properties (does not change with the amount of the sample) and extensive properties (changes with the amount of substance present).

Intensive properties include boiling/melting points, density, and conductivity. Extensive properties include mass, volume, and length.

Topic 5.4: Characteristic Properties

Characteristic properties are intensive properties which can identify a substance. These can include color, melting/boiling point, density at standard temperature/pressure, luster, and malleability or ductility.

Topic 5.5: Chemical and Physical Changes

Analogous to chemical and physical properties, changes can also be divided into two categories. If chemical properties of a substance change, a chemical change has occurred. However, if only physical properties of a substance change, a physical change has occurred.

Some possible indicators of a chemical change (meaning that these indicators do not conclusively show that a given change is chemical) include color change, solid disappearance, gas formation, precipitate formation, and light and heat production.

Physical changes are of the sort of melting, boiling, cutting, and bending. Chemical changes are of the sort of burning, cooking, water reactions, and air reactions.

Unit III: Chemical Bonds and Compounds

Lesson 1: Ionic Bonding

Topic 1.1: Ionic Bonding

Ionic bonding is a relatively simple type of bond where the majority of electrons stay in the valence of one atom. This definition defines the restriction of difference in electron negativity: any two elements can form an ionic bond if the difference of their electron negativities > 1.7 . The charge of the ions of a given element also make a difference. In addition, only elements with different ion charges (for example, Al^{+3} can form a bond with O^{-2} , but not Ca^{+2}).

For ionic bonding, note that this final requirement manifests in metal and nonmetal atoms forming ionic bonds with the metal becoming a cation (positive) and the nonmetal becoming an anion (negative). This happens because the electrons transfer from the metal to the nonmetal for filling valence shells, holding the atoms together with electromagnetism.

Topic 1.2: Crystals

Ironic compounds form crystals instead of molecules. A crystal is "a solid in which the particles are arranged in a regular, repeating pattern." The crystal lattice is the structure of this given crystal (how exactly the atoms/ions alternate throughout the crystal). Because ionic bonds don't form molecules, their chemical formula is a "formula unit," the smallest unit of the compound in the crystal. This is always the simplest ratio. For example, there will never be a Na_2Cl_2 because this would simplify to $NaCl$.

Topic 1.3: Polyatomic Ions

Ions are typically atoms with an extra electron, but ions can be made of multiple atoms (polyatomic ions). These ions act as single units and are always bonded together covalently. The charge is distributed across the entire molecule because the unit is covalently bonded and thus shares electrons. This can also cause some notational quirks — mostly that chemical formulas such as $Al_2(CO_3)_3$ can exist. Polyatomic ions include nitrites, nitrates, sulfite, carbonate, ammonium, cyanide, and numerous others. The traditional algebraic calculation of ion counts still applies in this case: for an ion with charge A and a second ion with charge B, they form an ionic bond of $A_B B_A$ (simplified to a ratio, and assuming A and B represent the respective elements).

Topic 1.4: The Formation of Crystal Lattices

Ionic bonds often form lattices because lattices have a lower potential energy than lone molecules, making it more stable and causing energy to be released when the crystal forms. Lattice energy describes this amount of energy. Lattice energy follows a few trends. It is greater for small ions such as lithium in comparison to sodium. It is also greater for ions with larger charges because both of these factors cause higher bond strength.

Topic 1.5: Properties Resultant from Lattice Energy

Higher lattice energies mean that a given crystal is more stable — causing high melting and boiling points. This also makes them hard, but it reduces solubility in water because — to dissolve — the ionic bonds have to be broken down. These ionic bonds also have very low conductivity as solids, but they have high conductivity as liquids or in a solution.

Lesson 2: Covalent Bonding

Topic 2.1: Orbital Overlap

Covalent bonds are different from ionic bonds in that electrons are a part of both atoms at the same time. These form exclusively between nonmetals. Similar to ionic bonds, covalent bonds occur because it is more stable to have a full outer shell of valence electrons—the octet rule. The H_2 molecule, for example, forms because they share their singular valence electrons with each other.

The number of valence electrons a given atom has can be determined as mentioned in earlier lessons. This number of valence electrons is often represented by dots surrounding the atom, especially when unbonded. The notation remains the same when bonded, except certain "bonding electrons" are shown on the shared side of both letters (it looks similar to " $H : F$ "). By convention, this single bond of two electrons is shortened to a singular line (like " $H - F$ "). The other valence electrons are referred to as "nonbonding electrons." The bonds

which share the electrons are referred to as "sigma bonds" if they are overlapping orbitals.

Topic 2.2: Equal and Unequal Sharing of Electrons

In all covalent bonds between two atoms of the same type, the electronegativity is precisely equal, so there is no charge difference across the bond. This is called a "nonpolar" bond. This can also occur with different atoms of similar electronegativities (difference ≤ 0.4).

Where electronegativity differs more significantly (≥ 0.5), the electrons are shared unequally and a polar bond develops. However, these polar bonds are not as strong as an ionic bond, being just a partial charge buildup.

Topic 2.3: Lewis Structures

The name of the notation used to show valence electrons are called Lewis Symbols. These have a chemical symbol (such as *Ne* or *B*) at the center with valence electrons drawn around it, spreading the valence electrons as equally as possible amongst the sides (if that's not possible, the uneven side(s) are unimportant). To generate a large molecule diagram, one draws a "Lewis Structure." These are created in the same way as simple covalent bond diagrams are (a trivial example of a Lewis Structure), by drawing bond lines (representing covalent bonds) between atoms after their Lewis Symbols have been noted. As a general rule, leftover electrons are next to the central atom when drawing Lewis structures.

Topic 2.4: Exceptions to the Octet Rule

While a full valence shell is typically always most stable, elements such as transition metals which can have more than just the 8 standard valence electrons, an expanded octet. These expanded octets have unique Lewis Structures, generally a central atom without valence electrons other than bonding electrons (still filling the octet by making another shell the new valence electrons).

Topic 2.5: Double and Triple Bonds

Another exception to previously mentioned Lewis Structure is double bonding. Sometimes, to fill atoms' octets, more than two electrons must be shared. O_2 is the clearest example of this – with 6 valence electrons, sharing two fills the octets of both. These are more stable than single bonds. N_2 has to use the logical extension of this, triple bonds, to fill their octets.

Unlike single bonds, double and triple bonds use more than sigma bonds. These use pi bonds, the overlap of p orbitals. These occur parallel to the current sigma bond, with one pi bond in a double bond and two pi bonds in a triple bond.

Topic 2.6: Resonance Structures

Resonance Structures are another exception to Lewis Structures. When multiple Lewis Structures could be possibly constructed from a given molecular formula (such as SO_2), these are both noted because experimental evidence suggests the real structure of molecules is a blend of the resonance structures.

Topic 2.7: Properties of Covalent Compounds

Covalent compounds are distinguishable from ionic compounds by several properties. These are composed of molecules, do not ionize in a solution, poorly conduct electrical charge, poorly conduct heat, and often have low melting and boiling points.

Lesson 3: Metallic Bonding

Topic 3.1: The Electron Sea

Metals have properties which cause unique characteristics:

- Large atoms
- Low electronegativity
- Low ionization energy

These properties create metals' "delocalized electrons," where they are not associated with a specific atom in a metal crystal but instead shared between nuclei and roaming throughout the crystal lattice. The simplest model which considers these is called the "electron sea model," where electrons are assumed to easily flow between nuclei and serve as "glue" for the metal atoms. This model describes metal bonding.

Topic 3.2: Molecular Orbitals

In metallic bonds, the orbitals that are overlapping share electrons, making "molecular orbitals" which determine the probable position of any given electron. The large number of atoms in a metal crystal means that there is a large number of molecular orbitals in the same crystal. These orbitals combine to form "bands," as predicted by band theory—another theory of metallic bonding which has more predictive power than the electron sea.

Topic 3.3: Resultant Properties

Because of metals' delocalized electrons, they have certain useful properties. These include high heat/electrical conductivity because the delocalized electrons can flow easily through the substance, and those same electrons can carry thermal energy. The high number of valence electrons and large atomic radii also contribute to the ease of movement of the delocalized electrons. Their other

unique properties of malleability, ductility, and luster can also be explained with their atomic characteristics. Their delocalized electrons make "flexible" bonds and thus allow the metal to be malleable and ductile instead of brittle. The band theory explains metals' lustrousness because the electrons move between bands (shared orbitals) when the metal absorbs light, which is successively released—causing the electrons to oscillate between bands. These properties are typically exacerbated by large numbers of valence electrons and large atomic radii, as mentioned earlier.

Topic 3.4: Alloys

Alloys are a type of mixture. Alloys are mixtures made by melting together metals and are thus homogenous. Properties of alloys are different from the properties of its composite metals, which is what makes them useful.

Alloys include brass (copper & zinc), rose gold (gold & copper), bronze (copper & tin or aluminum), and steel (mostly iron, chromium, & nickel).

Lesson 4: Intermolecular Forces

Topic 4.1: Intramolecular vs Intermolecular Forces

Intramolecular forces are, by definition, acting within a molecule. These are chemical bonds (covalent and ionic). Intermolecular forces act between molecules in a substance, attractively or repulsively. These can affect interactions between molecules and have 3 classes: hydrogen bonds, London dispersion forces, and Van der Waals forces. Intermolecular forces are significantly weaker than intramolecular forces.

Topic 4.2: Hydrogen Bonds

A hydrogen bond is "the attraction of a single hydrogen between a highly electronegative atom and another highly electronegative atom in a different molecule or a different region of a large molecule. These are commonly seen in water because of the strongly electronegative nature of the oxygen. It pulls electrons away from the hydrogen, creating a charge differential which acts like a hydrogen bond between hydrogens and oxygens of different molecules. Note that hydrogen bonding only occurs when hydrogens are directly bonded to highly electronegative atoms.

Hydrogen bonding affects the properties of substances which affect significantly because it is only roughly $\frac{1}{10}$ the strength of a real bond. The properties which it creates (when present) include higher melting and boiling points, high surface tension, high viscosity, highly structured solid forms (less dense than liquid form), and in some cases (such as water) polarity. This polarity can lead to macroscopic solubility properties. Ethane, a nonpolar, non hydrogen bonding chemical can only dissolve nonpolar covalent substances in contrast to water being able to dissolve ionic and polar covalent substances readily.

These are very common in organic molecules because oxygen and hydrogen are often present as major building blocks.

Topic 4.3: Van der Waals Forces

Van der Waals forces, unlike hydrogen bonds, occur in all atoms. These are the attractive or repulsive forces created by uneven distribution of electrical charge. This uneven distribution can occur from random variation, and it can occur from differences in atoms' electronegativities. Between polar molecules, a Van der Waals force is present, attracting the negative pole of one atom to the positive pole of the other. This is called a dipole-dipole interaction. Another common Van der Waals force is dipole-induced dipole interaction between polar and nonpolar molecules. The positive or negative end of the polar molecule pushes the same force away (negative repels electrons, positive attracts electrons), pulling them together. This induced force is weaker than a dipole-dipole interaction. The third Van der Waals interaction is called a London dispersion force. This is an induced dipole-induced dipole interaction. While it occurs in all atoms, it is only significant in nonpolar-nonpolar interactions because of how weak it is. London dispersion forces cause molecules to attract.

Topic 4.4: Applications of Intermolecular Forces

As these intermolecular forces fundamentally change the properties of materials, they can be used to stick to surfaces (such as in geckos' feet), or design soap which disrupts water's intermolecular forces giving it a high surface tension, melting point, and density.

Lesson 5: Nomenclature of Ionic Compounds

Topic 5.1: Charge of Ions

When forming ionic compounds, one of the atoms will necessarily lose electron(s) to the other, and because this usually aims for a full valence shell the number is typically the same between different bonds. The number which it loses or gains is called its oxidation state, and varies between atoms. Metals' oxidations are +1 for group 1, +2 for group 2, +3, and multiple, variant oxidation states for transition metals. For nonmetals in group 16 and 17, the oxidation states are -2 and -1 respectively.

Ionic compounds are neutral, allowing for algebraic determination of ion ratios/quantities. For example, if element A has an oxidation state of A and element B has an oxidation state of B, $Ax + By = 0$ where x is the number of A ions and B is the number of B ions. This works consistently.

Topic 5.2: Nomenclature of Ionic Compounds

Nomenclature is the rules for naming compounds, as developed by the International Union of Pure and Applied Chemistry (IUPAC). For ionic compounds,

it is always [cation metal] + [anion nonmetal modified with -ide]. This can also be performed in reverse: if one takes the name of an ionic compound, the component ions can be determined and from that the elements, ratio, and formula in that order.

Topic 5.3: Polyatomic Ions

This is the same as monoatomic ionic compounds, but without the -ide modification for anions. For example, $(NH_4)_2CO_3$ is ammonium carbonate.

Topic 5.4: Transition Metals

As mentioned earlier, transition metals are an exception to the singular oxidation states, and can each form multiple ions with different charges because the d and f orbitals allow different stable electron configurations. These have no easily discernible pattern, so it is easiest to memorize or look up oxidation states of a given transition metal. When converting chemical formulas with transition metals to the name of that ion, one uses roman numerals to signify the oxidation state. For example, in MnO , the Mn has a charge of +2 because oxygen always has an oxidation state of -2 and ionic compounds have 0 net charge, this meaning that the name is “magnesium (II) oxide.” Note that roman numerals ought be used only when the metal ions have multiple common charges.

Lesson 6: Nomenclature of Covalent Compounds

Topic 6.1: Nomenclature for Covalent Compounds

The element further left on the periodic table is listed first (e.g. CO_2 is carbon dioxide, not oxygen carbide). The second element ends with -ide (e.g. carbon dioxide). Numerical prefixes are used to indicate the number of atoms of a given element (1 of the second element=mono, 2=di, 3=tri, 4=tetra, 5=penta, 6=hex) such as in carbon tetrachloride or dinitrogen tetroxide.

Topic 6.2: Exceptions

The most major exception to these rules are traditional names such as water instead of dihydrogen monoxide, where the traditional name would be preferred. The other exception are compounds with hydrogen. In hydrogen compounds, a prefix for the amount of hydrogen isn't used such as in hydrogen sulfide (H_2S) or hydrogen sulfate (H_2SO_4). However, these chemicals are often acidic, which means they follow unique naming rules.

Topic 6.3: Acids and Bases

Acids increase the concentration of H^+ ions in a water (aqueous) solution. In the same solution, bases increase the concentration of OH^- ions. Both can be either ionic or covalent, but they have significant structural differences: acids

generally contain H atoms bonded to other atoms or polyatomic ions and ionic bases contain OH^- in their formula. Note that covalent bases do not contain OH^- . Acids turn blue litmus paper red while bases turn red litmus paper blue. Tastewise, acids taste sour and bases taste bitter. Bases also feel slippery or soapy.

Acids have two different methods by which they are named, depending on what type of acid they are. Binary acids (acids with exactly two atoms) are *hydro-[nonmetal]-ic*. Oxyacids (acids with polyatomic ions, in turn with one or more oxygen atoms) change dependent on the type of polyatomic ion. *-ate* ions form *-ic* acids and *-ite* anions form *-ous* acids (e.g. nitrate *-ic*, nitric acid and nitrite *-ous*, nitrous acid).

Bases' nomenclature is slightly less formulaic. Most covalent bases are amines, meaning that they are similar in structure to ammonia, but instead of a hydrogen, an R-group (a carbon or string of carbons) is present. The number of R-groups changes the name, such as primary amines (1 r-group) or tertiary amines (3 r-group). A hydrogen ion can attach to the Nitrogen's extra pair of electrons. Ionic bases follow the same rules as other ionic compounds.

Lesson 7: Polymers

Topic 7.1: Formation of Polymers

Monomers are the basic units of polymers—simple molecules like ethylene which can combine with each other by covalent bonding. Monomers share electrons (such as deconstructing a double bond to form a single bond to another atom) in order to form these bonds. Because they can each form covalent bonds with more than one other monomer, these can occur as repeating subunits of a large molecule (polymer). The process of forming this large molecule (also known as a macromolecule) is called polymerization. The properties of a given polymer are, notably, dissimilar from its constituent monomers. Ethylene, for example, is a colorless, flammable gas with a sweet taste and odor, but as part of a polymer, this forms the plastic of water bottles.

Topic 7.2: Polymer Classification

There are two types of polymers: natural and synthetic. These are precisely as they sound, made in nature and manmade respectively. Both are immensely useful in technological applications. For natural polymers, these make gloves rubber bands (from natural rubber), gloves, scarves, carpets (from wool), ties, and silk fabrics (silkworm cocoons). DNA (carries genetic information), starch (food and paper-making), and cellulose (plant stems, and used as thread such as cotton jeans) are three more incredibly common natural polymers.

Synthetic polymers uses include: polystyrene foam (makes take-out containers, packaging materials, egg cartons, and insulation), nylon (used to make ropes, nets, backpacks, cooking utensils, and stockings), polyethylene (plastic bags, bottles, toys), and vulcanized rubber (tires, shoe soles). The reason that

these are so useful in contrast to natural polymers is because they are flexible, hard, lightweight, strong, resistant, and often extremely cheap to produce. However, these are often not biodegradable, causing the synthetic polymers to build up in landfills as recycling is expensive. Also, the materials used to make synthetic polymers aren't always available.

Lesson 8: Properties and Uses of Saturated Hydrocarbons

Topic 8.1: Isomers and Hydrocarbons

Isomers are chemicals which are similar in one of two properties (if they are similar in both, they are the same chemical). These are structural isomers (same chemical formula) and geometric isomers (the structure looks the same, but the numbers are different). Hydrocarbons are chemicals with exclusively hydrogens and carbons, and saturated hydrocarbons are hydrocarbons with only single bonds. These are also known as alkanes. In contrast, unsaturated hydrocarbons contain at least one double or triple bond.

Topic 8.2: Hydrocarbon Notation

Hydrocarbons and organic compounds in general are often represented with standard Lewis structures, but there are some notational abbreviations which can be taken when drawing the chemicals. The first of these is abbreviation of methyl groups (CH_3 , CH_2 and CH ; note that CH_4 is never present in hydrocarbons with more than a single CH_4 because all of its atoms have full valence shells, leaving no more room for bonding). These methyl groups can also be written in opposite chemical order such as H_3C for convenience. The other notational abbreviation is a methyl group being written as a "joint" (no C , just connecting two lines or being the end of a line). Note that when this last notational trick is used in conjunction with double or triple bonds, the extra line(s) are written adjacent to where they would normally connect, not touching anything.

Topic 8.3: Straight Chains, Branched Chains, and Cycloalkanes

All compounds listed in this lesson are alkanes. Straight chains are the simplest type of alkane, as a singular straight line of carbons (with attached hydrogens). Branched chains are slightly more complex, with more than 2 "ends" (i.e. carbons attached to only one other carbon). Cycloalkanes have no "ends," forming rings of carbon.

Topic 8.4: Properties

These are generally:

- Fairly unreactive
- Combustible

- Insoluble in water but soluble in ether and other organic solvents
- having a higher boiling point with increased molecular mass.

Topic 8.5: Alkane Uses

Alkanes have several uses, often in mechanical applications. These include gasoline (pentane and octane), natural gas (methane and ethane), lubricating oil (carbon lengths of 17-35 carbons), and asphalt (carbon lengths > 35 carbons)

Topic 8.6: Nomenclature

For the 10 smallest straight chains, the naming system is a simple [*prefix*]-*ane*. For straight chains with 5 or greater carbons, the prefix follows Greek numerical prefixes, but the first 4 are methane, ethane, propane, and butane in that order. Branch chains have a multi-step naming system: one starts by naming the longest chain in the molecule and naming that (i.e. nonane or methane). Then, the substituents (chains coming off of the main chain) are identified and labeled (such as methyl groups). Next, all carbons on the long chain are numbered sequentially—starting at the end with the closest branch. After that, substituent groups are listed alphabetically (e.g. methyl before nitrogen before oxygen), and included in the name based on the location of the substituent groups. This becomes a molecule name such as 3,6-dimethylnonane if two methyls are attached to carbons 3 and 6. Cycloalkanes have a similar naming system, but with a couple of exceptions. Instead of starting the counting at the end, the counting starts at the longest substituent (because there is no end) and continues in the direction of the closest substituent. Numbering continues otherwise as normal, using standard cycloalkane names such as cyclopentane or cyclobutane, unless there is only one substituent. In the case of one substituent, no numbering is used. The name of the substituent is simply prefixed.

Lesson 9: Properties and Uses of Unsaturated Hydrocarbons

Topic 9.1: Alkenes, Alkynes

Alkenes are hydrocarbons with at least one double bond between carbon atoms. The chemical formula has the form C_nH_{2n} assuming only one double bond. Their naming convention is the same as alkanes except with the suffix "-ene." Alkynes have at least one triple bond between carbon atoms. They use the form C_nH_{2n-2} , and their naming convention is the same as alkanes except with "-yne." Both have a number prefixed to indicate where the double/triple bond is if the chain is long. These can be cyclic.

Topic 9.2: Cis Isomers and Trans Isomers

Because double and triple bonds use pi bonds, the molecules don't have freedom of movement about these bonds, meaning that there are two possible positions in which the adjacent carbons can occur relative to each other. These are cis isomers (the next carbon is in the same direction on both sides) and trans isomers (the next carbons are in opposite directions). Trans isomers are relatively uncommon. These are prefixed to compounds for identification

Topic 9.3: Aromatic Hydrocarbons

These are the double-bonded analog of cyclohexanes. The aromatic hydrocarbons always have 6 bonds in their basic unit (benzene), which alternates between double and single bonds. These are flat molecules, and can form into structures with multiple benzene units (such as naphthalene).

Topic 9.4: Boiling Points and Vapor Pressure

The boiling points of alkenes are lower than alkynes, and for both, as the boiling point increases with chain length. As mentioned in earlier lessons, vapor pressure is the pressure exerted by a vapor when the liquid and vapor are in equilibrium at a given temperature. For alkenes, vapor pressure decreases with the size of the molecule increasing, and like most substances increases with temperature.

Topic 9.5: Properties of Unsaturated Hydrocarbons

Almost all are combustible and insoluble in water. Alkynes have typically higher boiling points than alkenes or aromatics. These are all less dense than water, and alkenes tend to be pH-neutral while alkynes are typically acidic.

Topic 9.6: Applications of Unsaturated Hydrocarbons

Unsaturated hydrocarbons are sometimes present in food such as in polyunsaturated margarine or aspirin and other analgesic medicines (they contain chemicals derived from aromatic hydrocarbons). Rubber is made up of alkene monomers, polystyrene is made of hydrocarbons, and acetylene is an alkyne used as fuel in welding torches, gas lamps, and electric generators.

Unit IV: Chemical Reactions and Limiting Reactants

Lesson 1: Evidence of Chemical Reactions

Topic 1.1: Chemical Changes and Chemical Reactions

Chemical reactions accompany all chemical changes, and are effectively equivalent. Chemical reactions, by definition, are a rearrangement of the atoms in reactants which result in a new substance (product). Oxidation as in rust, combustion, and rising bread are all chemical reactions. An arrow (\longrightarrow) is often used to represent the change from listed reactants to listed products.

Topic 1.2: Rate of a Chemical Reaction

Chemical reactions can occur over a short period of time (fast) such as a firework, or they can take a long period of time (slow), such as limestone decaying. The speed at which this occurs is called the rate.

Topic 1.3: Indicators/Evidence of a Chemical Reaction

As noted in previous lessons, when a chemical reaction occurs, several things can happen which are common between chemical reactions. These include gas formation, solid (precipitate) formation, a color change, and always some form of energy change. The energy change can change the temperature, a decrease if the reaction is endothermic and an increase if the reaction is exothermic. Light is also emitted as another signal of energy change.

These indicators do not prove a chemical reaction as a hard and fast rule. Physical changes can be accompanied by gas formation (mentos and coke), solid formation (freezing), temperature change (boiling water), and color change (mixing dyes). Generally, the more indicators of a chemical reaction which are present, the more likely a chemical reaction has occurred.

Lesson 2: Writing and Balancing Chemical Equations

Topic 2.1: Equations

Equations represent a given chemical reaction (changing listed reactants on the left to listed products on the right, separated by pluses in both cases). These equations can represent each substance with words (such as water or sodium bicarbonate) or chemical formulae (such as $H_2O(l)$ or $12NaHCO_3(g)$). In special cases, a "model" (pictographic representation) of the formula can be used.

The formulae are the most common and accepted. The letters next to the chemicals' formulae represent the state (s is solid, l is liquid, g is gas, aq is aqueous). The numbers before the chemicals represent "counts" (ratios) of

moles/molecules and allow balancing by using algebra and the principle of conservation of matter (if O_2 is at left and H_2O is at the right, with no other oxygens, there will be the same number of oxygens on each side (double at the right)). The smallest possible numbers are preferable.

Often, reaction conditions are added above the arrow or the arrow becomes bidirectional to symbolize a reversible reaction. The reaction conditions are mostly self explanatory and there are 4 common symbols: a Δ for a heated reaction, 25° for a reaction at 25° , $2atm$ for a reaction at 2 atmospheres, and Pt for platinum as a catalyst.

Lesson 3: Percent Composition and Molecular Formula

Topic 3.1: Mole Ratios

A mole is a specific number of a given particle (ion, atom, molecule, etc.). The number is Avogadro's number and is chosen so that the mass of a mol of ^{12}C is 12g. Because this isn't reliant on mass, just the amount of particles every chemical has a molar ratio between its composite elements easily determinable from the empirical formula. CO_2 , for example, has a mole ratio of $C : O = 1 : 2$. Note that this isn't always just the numbers of elements listed in the formulae. N_2O_4 has a mole ratio of $N : O = 2 : 4 = 1 : 2$, with $1 : 2$ preferred as the simplest ratio and $2 : 4$ as the true ratio. These ratios also need not always include every element. For ammonium nitrate (NH_4NO_3), the $H : O$ ratio is $4 : 3$.

Topic 3.2: Empirical and Molecular Formulae

The true ratios and mole ratios mentioned earlier generate molecular and empirical formulae respectively (e.g. a given ratio of elements such as $C : H = 1 : 2$ becomes a formula such as CH_2). This implies that all empirical formulae have the simplest ratio of their elements, but molecular formulae can have any multiple of a simple ratio.

Topic 3.3: Percent Composition

In order to accurately analyse compounds, it can be useful to know the proportions of masses of an element. This is not computationally difficult as the percent composition of a given element in a given compound is the mass of the element in the compound divided by the mass of the compound. Note that for determining this ratio molecular and empirical ratios are equally well-suited. These ratios can also be determined for more complex chemicals than just elements. In hydrates (crystals which integrate water in their lattice), one can determine the percent composition of water by an identical calculus. Note that in a hydrate, the salt forming the crystal is called anhydrous when the crystal doesn't have water.

Lesson 4: Types of Reactions

Topic 4.1: Types of Chemical Reactions

There are five main types of chemical reactions: synthesis (2+ reactants combine to form a single product), decomposition (a single compound breaks down to form 2+ new substances), combustion (reaction of a substance with oxygen—includes oxidation), single replacement (reaction in which one ion displaces another to form a new compound), and double replacement (two ionic compounds exchange ions to form new products)

Topic 4.2: Differentiating Single and Double Replacement Reactions

In single replacement, both the reactants and products have one element and one compound, where the reactant element becomes part of the product compound. In double replacement, reactants and products are ionic compounds and the same ions are present—just reordered—between the reactants and products.

Topic 4.3: Activity Series

The activity series describes how reactive elements are compared to each other. For example, lithium is much more reactive than gold. This can allow us to make predictions about if a reaction will take place or not. For example, we know that $AgNO_3 + Cu$ are more likely to react than $Cu(NO_3)_2 + Ag$ because copper in its pure form is more reactive than silver is in its pure form. This generalises. Note that halogens' and metals' activity series are different. Metals' activity series is long, but the halogens' activity series is short ($F_2 > Cl_2 > Br_2 > I_2$).

Lesson 5: Limiting Reactant and Percent Yield

Topic 5.1: Limiting Reactants and Stoichiometry

Stoichiometry is the relationship between the relative quantities of substances taking part in a reaction or forming a compound, typically a ratio of whole integers. Stoichiometry is typically used in chemistry to determine the ratios between substances in which they react. For example, H_2 and N_2 react to form NH_3 in a ratio of $H : N = 3 : 1$. However, if there are 5mol of H_2 and 1mol of N_2 , then there is left 2mol of H_2 , called the excess reactant. Because more nitrogen would allow continued reaction, nitrogen is called the limiting reactant. In general, the limiting and excess reactants in a reaction with uneven proportions can be determined with a similar calculus. Stoichiometry also extends to comparisons between reactants and products. If $2O_2 + H_2O \rightarrow H_2O_5$, 10 fully reacted moles of oxygen form 5 moles of the substance on the right (note that the substance makes no sense and isn't real).

Topic 5.2: Theoretical Yield and Percent Yield

Theoretical yield is “the ideal maximum amount of a product that can be produced during a reaction, calculated from stoichiometric relationships.” All examples above assumed that reactions occur perfectly, consuming all of the reactants and creating the correct amount of product and nothing else. The amount of product created was the theoretical yield. However, real life doesn’t usually have the same yield (actual yield) as predicted. Actual yield is almost always less than theoretical yield due to suboptimal reactant conditions and random factors. Percent yield is the ratio between actual yield and theoretical yield.

Unit V: Stoichiometry and the Gas Laws

Lesson 1: Molar Masses

Topic 1.1: Moles and Avogadro’s Number

Moles (mol) are the SI unit for the amount of a substance, one mol being the same number of particles of a given substance as there are atoms in 12g of C-12, or 6.02×10^{23} . Avogadro’s number is that number of particles in a mol, and it is calculable from basic units of a carbon-12 atom, namely its atomic mass of $12amu$ and the fact that $1amu = 1.66 \times 10^{-24}g$

Topic 1.2: Molar Mass and Average Atomic Mass

The mole and Avogadro’s number relate average atomic mass and molar mass. Average atomic mass is the statistic displayed on the periodic table below a given element, in $\frac{amu}{atom}$. This is extensible to any arbitrary molecule, still in units $\frac{amu}{particle}$. Note that for a given substance, the molar mass (units $\frac{g}{mol}$) is equal to the average particle mass by definition of the mol.

Lesson 2: Introduction to Stoichiometry

Topic 2.1: Stoichiometry and Ratios

Stoichiometry is based on ratios, such as *Reactant : Product*, *Reactant : Reactant*, or *Product : Product*. In the previous lesson, this was explored as the way to determine theoretical yield, so the foundation should already be set. Note that these ratios are molar ratios calculated from the ratios of the coefficients next to each compound in a balanced chemical equation. These ratios represent the amount of any two given substances in a balanced chemical reaction, ignoring total amount, and can be used to calculate the amounts of byproducts, amount of reactant required to react completely, limiting reactants, storage requirements, and how much product is created with a certain amount of reactant.

Lesson 3: Stoichiometric Calculations

Topic 3.1: Molar Mass and Mass-to-Mole Ratio

Molar mass is in units of grams of element per mol. For example, $\frac{58.44gNaCl}{molNaCl}$. This can be expanded to a given number of moles or a mass using dimensional analysis (note that dimensional analysis permits inversion of a ratio such as $\frac{molNaCl}{58.44gNaCl}$). Also, for clarity's sake, the element formula next to each unit is necessary, and doesn't cancel.

Topic 3.2: Interchemical Calculations

Because of stoichiometric ratios as mentioned in the previous lesson, such as $\frac{molNaCl}{molCl^+}$ in a table salt decomposition reaction, dimensional analysis can be further used to determine the mass of product created by a given chemical reaction.

Lesson 4: Gas Laws

Topic 4.1: Boyle's Law

All gas laws are merely extensions of the ideal gas law mentioned in Unit II Topic 1.1. The ideal gas law is $\frac{P_1V_1}{N_1T_1} = \frac{P_2V_2}{N_2T_2}$ (N is number in moles, T is temperature in Kelvin, P is pressure, V is volume, and this holds for any two ideal gases according to Avogadro's Law). The first of these derivative laws is Boyle's Law: where quantity and temperature remain constant, pressure is inversely related to volume. This allows simple calculations relating changing volume and/or pressure of a gas.

Topic 4.2: Other Laws

Law	Relation
Boyle's Law	Volume and Pressure are inversely related
Charle's Law	Volume and Temperature are directly related
Gay-Lussac's Law	Pressure and Temperature are directly related

Topic 4.3: Partial Pressure

Partial pressure is "the fraction of the total pressure exerted by a mix of gases that is contributed by an individual gas." (i.e. the pressure of one gas in a mixture). $Partial\ pressure = \frac{Pressure\ if\ there\ were\ only\ the\ one\ gas}{Actual\ pressure}$. Dalton's Law states that the total pressure is equal to the sum of the partial pressures. Note that the fraction which the partial pressure makes up of the total pressure is equal to the molar fraction by the ideal gas law and the fact that temperature and volume are the same for all gases.

Lesson 5: The Ideal Gas Law

Topic 5.1: Avogadro's Law

Avogadro's Law states that "the volume of a gas is proportional to the moles of the gas when pressure and temperature are kept constant." This is the N term in the ideal gas law mentioned in last lesson. Note that this is unrelated to the gas's makeup.

Topic 5.2: Derivation of the Ideal Gas Law

From Boyle's Law ($V \propto \frac{1}{P}$), Charle's Law ($V \propto T$), and Avogadro's Law ($V \propto n$), the ideal gas law of $V \propto \frac{nT}{P}$ can be derived. This is equivalent to $V = R(\frac{nT}{P})$ or $PV = nRT$ where R is the gas constant. Because the ideal gas law is independent from the gas's identity, R has only one value. However, ideal gases can diverge from real gases significantly in high pressure or low temperature environments.

Lesson 6: Gas Stoichiometry

Topic 6.1: Implications of Molar Volume: Avogadro's Principle

As stated repeatedly in the previous two lessons, the volume of a mole of gas is irrelevant to its composition. This is exemplified by molar volume: one mole of any ideal gas at standard temperature and pressure ($0^\circ C = 273K$ and $1atm = 101.3kPa = 760torr$) always has a molar volume of $22.4L$. Molar volume varies with temperature and pressure but is consistent between substances. Molar volume implies Avogadro's Principle: "if two gas samples contain the same number of particles, they will have the same volume at a given temperature and pressure." Avogadro's Principle implies that in stoichiometry (ratios of coefficients in chemical reactions), the numbers can represent ratios of particles, moles, or volumes.

Unit VI: Reaction Rates and Equilibrium

Lesson 1: Reaction Rate

Topic 1.1: Energy Diagrams and ΔG

G is the free energy of a system, how much energy is readily available to do work in a system. ΔG is the change in the free energy over time, and ΔG_{rxn} is the change in free energy during a reaction. If $\Delta G_{rxn} < 0$, then the reactants have more free energy than the products, and this type of reaction occurs "spontaneously:" if any parts of the chemicals meet the activation energy for the reaction, the reaction is able to continue. The opposite of this type of reaction is nonspontaneous reactions where $\Delta G_{rxn} > 0$. This means that the products have more free energy than the reactants, and energy has to be constantly added to make the reaction continue.

Topic 1.2: Activation Energy

Activation energy (E_A) is the minimum amount of energy to initiate a chemical reaction (this is brought on by an increase in free energy of compounds during the reaction—meaning that energy is absorbed by the reaction before the “hump” of the activation energy is reached). This is a barrier to the reaction and must be overcome before the reaction will proceed. For all stable compounds, the reaction energy is positive because otherwise they are unstable and decompose readily (the higher the activation energy, the less likely the reaction is to occur and thus more slowly it occurs).

Topic 1.3: Chemical Bonds

Chemical reactions rearrange atoms in molecules by making and/or breaking bonds. Bonds, as mentioned in previous units, are stable electron configurations—implying that energy must be input to start decomposing or making bonds. Note that energy still must be input to start most spontaneous reactions, but they will continue on their own because energy from previous reactions supplies activation energy for other reactant molecules.

Topic 1.4: Reaction Rate

Reaction rate is how quickly reactants convert to products. This depends inversely on activation energy. High activation energies (like in graphite to diamond reaction) mean low reaction rates, and low activation energies (like baking soda and vinegar) mean very high reaction rates. This is completely unrelated to ΔG_{rxn} .

Topic 1.5: Collision Theory

Collision Theory is the model that explains chemical reactions which assumes that chemical reactions' rates are proportional to their collision rate (note that, in reality, orientation and speed of the molecules are important, but the rate is still proportional to the number of collisions). Kinetic molecular theory allows the idea of a Maxwell-Boltzmann Distribution in that molecules have different speeds even at the same temperature. The Maxwell-Boltzmann Distribution describes how many molecules are at a given speed for a given temperature. When graphed, the vertical axis typically represents the count/frequency of particles and the horizontal axis the velocity of the particles. These distributions are typically bell curves, and the “longer” the curve, the higher the temperature

Topic 1.6: Effective Collisions

Effective Collisions, as mentioned in the previous topic, are necessary (as compared to simply collisions where all possible touching of molecules results in a reaction) for chemical reactions to occur. What contrasts these effective

collisions is a high enough kinetic energy and the reactants' facing the correct directions to exchange atoms.

Topic 1.7: Effect of Concentration on Reaction Rate

Similarly to temperature increasing collision rate by increasing molecule movement and pressure, increased concentration causes more frequent collisions—increasing the rate of reaction (concentration meaning number of molecules per volume in an aqueous solution, for example).

Topic 1.8: Effect of Pressure on Reaction Rate

There are 2 other ways to increase pressure other than temperature: increasing the amount of reactants and decreasing volume (which makes sense with the ideal gas law). However, note that increasing the pressure is only useful with reactions exclusively between gases because liquids and solids don't become more dense with higher pressures.

Topic 1.9: Effects of Surface Area on Reaction Rate

Reactions occur at the surface of solids, so decreasing the particle size (such as from one large cube to 8 smaller cubes) will increase the reaction rate. This can also apply to changing the shape (a sphere will have a lower reaction rate than a spiky ball) of the solid reactants. However, the particles are typically too small as for this to be realistic.

Lesson 2: Reaction Pathways

Topic 2.1: Reaction Pathway Graphs

A reaction pathway graph is a diagram of the change in energy between reactants and products. This is sometimes called a potential energy diagram because it graphs the Gibbs Free Energy/potential energy over time during a reaction. Energy is typically on the vertical axis, and time/reaction progression is on the horizontal axis (meaning that reactants are to the left and products to the right).

Topic 2.2: Exothermic Reactions

Exothermic Reactions are reactions which release energy. This occurs when the potential energy of the reactants is higher than the potential energy of the products (kinetic energy is released to conserve energy). These include condensation, combustion, dissociation of strong acids, solidification of cement, concrete, and epoxy, and the thermite reaction. These follow the exothermic reaction pathway which is the decrease in potential energy of the materials.

Topic 2.3: Endothermic Reaction Pathway

Endothermic reactions are the opposite of exothermic reactions: they absorb energy. These include melting, photosynthesis, decomposition and dehydration reactions, and the dissociation of some salts. In these, reactants always have lower potential energy than products. For both endothermic and exothermic reactions, the reaction has the property ΔH_{rxn} — the change in potential energy of the substances. Endothermic reactions have positive ΔH because the products have higher potential energy than the reactants. The opposite is true of exothermic reactions.

Topic 2.4: Activation Energy and the Activated Complex

As mentioned in previous lessons, all reactions involving stable compounds have an activation energy—the minimum amount of energy needed to start a chemical reaction (such as heating a match to make it start burning). The activated complex is a “short-lived, high-energy intermediate between reactants and products,” which is necessary to complete a reaction and the reason why activation energy is required to bring substances together to form the activated complex. The activation energy is defined as the difference between the activated complex’s potential energy and the reactants’ potential energy. This is represented as E_a . As activation energy increases, reaction rate decreases.

Lesson 3: Catalysts

Topic 3.1: Catalysts

Catalysts are substances which increase the rate of a reaction but aren’t consumed by the reaction. This is by the creation of an alternate reaction pathway with lower activation energy. As noted in a previous lesson about notation, chemical equations with catalysts have the catalyst written above the arrow.

Topic 3.2: How Catalysts Work

Catalysts decrease the activation energy/increase reaction rate by holding reactants in the right orientation to interact to increase the number of effective collisions. They can also form temporary intermediate substances for a different reaction pathway, and they sometimes weaken or break reactant bonds.

Topic 3.3: Heterogenous Catalysts

Heterogenous catalysts are any catalyst which is in a different state of matter than the reactants—making it easier to remove when the reaction completes. These include platinum in a catalytic converter or iron in ammonia production.

Topic 3.4: Enzymes

Enzymes are complex proteins which are produced biologically and catalyze biochemical reactions at body temperature. Most are specific to particular reactions and substances. Enzymes function by interacting with a substrate (reactants) at the active site, forming a sort of lock and key structure called the enzyme-substrate complex. Then, the reaction occurs and the products are formed. Enzymes are sometimes inhibited by body conditions (different temperatures or acidity can deform proteins) and/or extra compounds interacting with it at the active site (blocking interaction with the actual substrate) or at an allosteric site (deforming the protein). The enzymes can form feedback loops with the body to regulate conditions and limit their deformation.

Lesson 4: Reversible Reactions and Equilibrium

Topic 4.1: Equilibrium in Reversible Reactions

Equilibrium is the state where a system is balanced (like boiling water in a closed vessel applying pressure to the liquid part until it stops boiling). A reversible reaction is a reaction where neither the forward (*reactants* \rightarrow *products* and *reactants* \leftarrow *products* can both occur). Note that reactants and products are always placed in that order regardless of the direction of the arrow. The favored reaction direction can change with environmental conditions like pressure and temperature, so equilibrium is possible with reversible reactions (manifesting as a mixture of reactants and products)

Topic 4.2: Forward and Reverse Reaction Rates

Reactions are constantly happening if the reactants are available and enough kinetic energy is present, so equilibrium with reversible reactions must be dynamic equilibrium—the system stays at the same concentration ratios overall, but any specific part is constantly changing. Forward reaction rates and reverse reaction rates change how much the mixture changes in one way or the other (i.e. are more reactions occurring *reactants* \rightarrow *products* or *reactants* \leftarrow *products*)

Topic 4.3: Chemical Equilibrium

Because chemical reactions' rates are reliant on concentration, equilibrium can occur even when a forward or reverse reaction rate is higher. If the forward reaction rate is higher where reactants and products are the same, then there are more products (products are favored), and the equilibrium lies “to the right” on the chemical equation. Similarly, if the reverse reaction rate is higher at a 50/50 mix, there are more reactants, the equilibrium lies “to the left,” and reactants are favored. Note that at equilibrium, the reverse reaction rate is always the same as the forward reaction rate by definition.

Topic 4.4: The Equilibrium Constant (K_{eq})

The Equilibrium Constant is $\frac{[C]^c[D]^d}{[A]^a[B]^b}$ for the reversible reaction $cA + cB \leftrightarrow cC + dD$ if $[K]$ is the concentration of substance K in terms of molarity for aqueous solutions and partial pressure for gases, and k is the coefficient of substance k in the chemical equation. Note that pure solids and pure liquids do not appear in equilibrium constant calculations. The equilibrium constant can typically be interpreted as: $K < 1$ means reactants are favored; $K > 1$ means products are favored.

Lesson 5: Shifts in Equilibrium

Topic 5.1: Changes in Reactant or Product Concentration

Stresses are external changes which disrupt chemical equilibrium. These include: change in pressure, change in temperature, and change in concentration. Le Chatelier's principle states that "if a stress is applied to a chemical system at equilibrium, the system will respond by shifting in a direction to counteract the stress and a new equilibrium will be established. In a chemical system, for example, if a concentration change occurs, then the forward or reverse reaction rate will increase enough to bring back the equilibrium. This is true of any change in chemical concentration: K_{eq} always remains constant as long as no other features of the environment vary.

Topic 5.2: Common Ion

A common ion is the ion which is in both a system's chemical equation (individually like H^+) and an external ionic compound (such as HCl). For example, adding hydrochloric acid (HCl) to a system with Chlorine as a reactant causes the system to respond by shifting the equilibrium further in favor of the forward reaction (to the right).

Topic 5.3: Changes in Pressure

Changes in pressure shift the equilibrium in gaseous mixtures. Equilibrium typically shifts in the direction of lower molecules (e.g. $A \rightarrow 2B$ would shift the equilibrium to the left if pressure increases).

Topic 5.4: Changes in Temperature

Changes in temperature also change the equilibrium because heat is either a reactant or a product (the reaction is necessarily endothermic or exothermic). Because heat acts like a substance on its respective side, an increase in temperature will move the equilibrium away from its position in the location (e.g. $A \leftrightarrow products + heat$ shifts to the left as temperature increases)

Topic 5.5: Reaction Quotient

For non-equilibrium concentrations, applying the equilibrium constant formula gives the reaction quotient, which helps to inform if the forward or reverse reaction is currently preferred. If $Q > K_{eq}$, the reverse is typically preferred. Conversely, if $Q < K_{eq}$, the forward is preferred.

Unit VII: Energy in Chemical Reactions

Lesson 1: Energy

Topic 1.1: Basic Definitions

Energy is “the ability to do work or cause change.” The SI unit for energy is the joule or the kilojoule. Other units have also been historically used to represent energy such as calories, kilocalories, BTUs, and kilowatt-hours. Kinetic energy is energy associated with movement. This is related to the amount of work done by an object, and its mass/velocity. Formally, a force consistently applied over a given distance does the work equal to $F * d$. Energy inputted to a system must always be greater than or equal to work done due to the conservation of energy and non-conservative forces like friction.

Topic 1.2: Potential Energy

Potential energy is any energy associated with the position of an object. The most common form is gravitational potential energy because dropping it can generate an equal or lesser amount of kinetic energy from whichever gravitational field it is in. If the field is approximately constant, then $G_{PE} = mgh$ where m is mass, g is gravitational attraction, and h is height.

Topic 1.3: Mechanical Energy

Mechanical Energy is the sum of all kinetic and potential energy of an object.

Topic 1.4: Thermal Energy

Thermal energy is the energy associated with temperature, which can flow as heat. This is dependent on mass and absolute temperature. Heat is “energy that flows from a warmer object or substance to a cooler object or substance.” Heat constantly flows until all objects in a system reach equilibrium.

Topic 1.5: Electromagnetic and Chemical Energy

Electromagnetic energy is “energy associated with electric fields, magnetic fields, and electromagnetic radiation.” This includes visible light (such as from the sun), X-rays, radio waves, or electricity (like the type which flows through an outlet). Chemical energy is a type of potential energy stored in chemical

bonds. Transfers of chemical energy occur in chemical reactions (endothermic = storing, exothermic = releasing energy).

Topic 1.6: Law of Conservation of Energy

Similar to the law of conservation of matter, the law of conservation energy states that "energy cannot be created or destroyed, but it can change from one form to another." These exchanges include, for example, burning gasoline to turn chemical energy into thermal energy. There are two types of systems when talking about the exchange of energy: open systems (can exchange matter and energy) and closed systems (can exchange energy but not matter)

Topic 1.7: Energy Transfer and Transformation

Energy transfer is the movement of energy in the same form. Heat transfer of the type of energy dispersing from a fire is energy transfer. Energy transformation, however, is the change in energy from one form to another like nuclear to thermal energy or electromagnetic to mechanical energy. Note that both follow the law of conservation of energy.

Topic 1.8: Sources of Energy

Energy can't be created or destroyed, so it must come from somewhere. For human use, the earth's interior, the sun, fossil fuels, radioactive elements (like U^{235}), moving air/water, and food can be considered sources of energy.

Lesson 2: Heat

Topic 2.1: Thermal Energy

Thermal Energy is the same as the kinetic energy of the molecules of a substance. This is a property of all matter. Thermal Energy can be measured with temperature—the average kinetic energy as opposed to the total kinetic energy of a substance. Temperature can be described with $^{\circ}F$, $^{\circ}C$, and K .

Topic 2.2: Heat and Conduction

Heat is "the transfer of kinetic energy between molecules as faster-moving molecules collide with slower-moving molecules. Heat always trends towards equilibrium and thus generally flows from high kinetic energy to low kinetic energy. Conduction is a type of heat flow which involves direct contact of two substances such as a metal spoon and a hot pot. This, like heat in general, tends towards equilibrium, "taking" energy from the fast molecules and "giving" it to the slow molecules.

Topic 2.3: Heat Flow during Chemical Processes

Exothermic and endothermic chemical reactions decrease and increase the potential energy of the involved substances, respectively. However, in order to maintain conservation of energy (total energy remains the same), the environment either increases its own heat (exothermic) or decreases (endothermic).

Lesson 3: Calorimetry

Topic 3.1: Specific Heat Capacity

Heat capacity is “the quantity of heat needed to raise the temperature of a given sample of a substance by one degree Celsius/Kelvin.” This is described in units of $\frac{\text{energy}}{\text{temperature}}$ (ex: $\frac{J}{^\circ C}$ or $\frac{cal}{K}$). Note that this is an extensive property because the mass of a sample is relevant (i.e. a larger sample takes more energy to heat by the same amount). Specific heat capacity (C_p) is the heat capacity per gram, so this has the units of $\frac{\text{energy}}{\text{mass} \cdot K}$. Specific heat can change with temperature, especially in gases. This is related to intermolecular forces. Heat absorbed or released (q) can be related, mathematically, as $q = mC_p\Delta T$ where m is mass and ΔT is the change in temperature.

Topic 3.2: Calorimetry Definition

Calorimetry is “the use of a calorimeter to measure energy given off or absorbed during a physical or chemical process,” where a calorimeter is a device designed to do exactly that.

Topic 3.3: Bomb Calorimetry

A bomb calorimeter is a specific type of calorimeter which doesn’t use a bath of water (solution) for the reaction. Instead, it has a separated vessel inside of a bath of water to measure more reactions. These are often used to measure combustion heat release.

Topic 3.4: The Calorie Content of Foods

Nutrition information labels give calorie counts (note that a food Calorie is a kilocalorie or $4.184kJ$). These are determined with bomb calorimeters by burning the food and determining heat released during the combustion reaction.

Lesson 4: Thermochemical Equations

Topic 4.1: Enthalpy

Bonds contain potential energy, and breaking or forming bonds can release or absorb energy. All reactants and products contain chemical energy. Enthalpy (H) measures the heat and internal energy in a system. Enthalpy in a chemical reaction is a state function, a type of function which depends only

on the beginning and end of the process (not the middle). This makes sense because enthalpy of reaction is the same if a catalyst is used or not. Enthalpy is also a linear function, so the only important part is the potential energy of the involved reactants and products. This potential energy is called enthalpy of formation (ΔH_f). It is defined relative to the zero-point of an element in a “standard state” at STP. Chlorine gas, for example has a ΔH_f of $0 \frac{kJ}{mol}$.

Topic 4.2: Enthalpy of Reaction

Enthalpy of a reaction (ΔH_{rxn}) is the energy absorbed (positive) or released (negative) during a chemical reaction. Hess’s Law of heat formation states that this property is linear with enthalpy of formation, i.e. the enthalpy of a reaction is the enthalpy of the reactants subtracted from the enthalpy of the products. A thermochemical equation is a chemical equation which includes this information and states of involved substances.

Topic 4.3: Manipulating Equations

If the direction of a chemical equation is reversed, ΔH_{rxn} becomes $-\Delta H_{rxn}$. If a chemical equation is multiplied by a coefficient k , ΔH_{rxn} becomes $k\Delta H_{rxn}$. Note that ΔH_{rxn} typically refers to integer molar amounts of compounds in a balanced chemical equation.

Topic 4.4: Enthalpy of Combustion

Combustion, another chemical reaction, also has a certain amount of enthalpy. Combustion reactions are typically very formulaic, so the energy of a given combustion reaction (such as burning propane) is known by the scientific community. Note, however, combustion enthalpy can still be determined with standard techniques.

Lesson 5: Enthalpy and Phase Changes

Topic 5.1: Heat and Phase Changes

Previously, heating and cooling curves have been described. These are graphs of a substance across temperature and added heat, showing its phase changes. While in different states of matter, substances have different specific heats. While changing between substances, temperature is constant, but energy is not. For melting, the energy to transform a mole of a substance at the melting point, but completely solid, is called the molar heat of fusion (ΔH_{fus}). The heat of fusion can also be given in terms of $\frac{J}{g}$ but is typically provided as $\frac{J}{mol}$. The energy released during the opposite process, freezing, is equal to $-\Delta H_{fus}$. Boiling/vaporizing follows a similar pattern: the molar heat of vaporization is ΔH_{vap} , and the molar heat of condensation (not a real measurement) is $-\Delta H_{vap}$. These can also be given in terms of mass.

Topic 5.2: Effects of Heat on Phase Changes

During phase changes, temperature doesn't change, so kinetic energy doesn't change. However, heat input has to increase some form of energy, so it increases potential energy—enough so that the movement of the particles overcomes the intermolecular forces. Similar to chemical reactions, potential energy increases during endothermic reactions and decreases during exothermic reactions.

Topic 5.3: Relationship between Enthalpy

$$\Delta H_{vap} > \Delta H_{fus} > 0$$

Topic 5.4: Sweat and Temperature Regulation

Because evaporation is endothermic, the human body uses sweat to cool itself by making sweat evaporate (temperature doesn't increase because no temperature changes occur during phase changes).

Lesson 6: Enthalpy of Reaction

Topic 6.1: Equation Combination

By Hess's Law of Heat Formation, chemical equations can be added like linear equations. This is most useful for combined processes (like writing photosynthesis from its component reactions). The goal of combining equations is to write a balanced equation of the products respective to reactants as chosen.

Topic 6.2: Diagramming Hess's Law

Because enthalpy is a state function, and the potential energy of a given compound or group of compounds is constant, a set of chemical equations can be diagrammed as a set of transformations from one set of compounds (i.e. $2C(s)+H_2(g)$ to $C_2H_2(g)$ is one transformation), and from that, the difference in the final product's potential energy and the reactants is the change in enthalpy. Note also that upwards arrows on the diagram are positive, and downward arrows are negative.

Topic 6.3: Reaction Energy Diagrams and Enthalpy Diagrams

Reaction energy diagrams and enthalpy diagrams are both ways to describe the change in potential energy of a chemical reaction. Enthalpy diagrams are more detailed in the process, but reaction energy diagrams describe reaction progress with more detail. Also, reaction energy diagrams only display one reaction at a time unlike enthalpy diagrams displaying the total and individual reactions. Both put energy on the vertical axis.

Unit VIII: Mixtures, Solutions, and Solubility

Lesson 1: Mixtures and Solutions

Topic 1.1: Heterogenous and Homogenous Mixtures

As described extensively in previous lessons, homogenous mixtures are approximately congruent between any two arbitrarily chosen volumes, but heterogenous mixtures are not.

Topic 1.2: Solution

A solution is a type of homogenous mixture which appear as one phase. It has two components: a solvent (the “dissolver”), and a solute (the “dissolved”).

Topic 1.3: Suspensions and Colloids

Suspensions are similar to solutions, but instead of being homogenous they are heterogenous mixtures which have large enough particles to settle out or be filtered out. This includes muddy rivers, milk, and blood. Colloids are a subset of suspensions with significantly smaller particles dispersed in a way that makes it difficult to easily filter or quickly settle.

A few features of suspensions, colloids, and suspensions can be used to differentiate them. The first of these is Brownian motion: the constant random motion of particles is observed in colloids and solutions but not in suspensions. Quick settling is observed in suspensions by definition but not in most colloids or any solutions.

Topic 1.4: The Tyndall Effect

The Tyndall Effect differentiates colloids from solutions because they can be easily confused. In colloids, light passing through is scattered by the particles like dust in air or flour in water. This scattering, the Tyndall Effect, notably does not occur in solutions. The scattering of light looks like a coherent beam of light.

Topic 1.5: Techniques to Separate Solutions

There are a few common techniques to separate solutions. These have been listed before, and include distillation (boiling liquids and collecting the vapor of liquids as they boil at different temperatures), crystallization (vaporizing the solvent to leave the solute(s)), and chromatography (separating solutes by density or particle size).

Colloids are slightly easier to separate. Centrifuges (spinning the colloid to separate solutes by density with centrifugal force), long standing (leaving the solute to settle out of the solvent), and boiling/heating (heating or running electricity through the colloid to coagulate—thicken to a solid or semisolid—it).

Lesson 2: Reactions in Aqueous Solutions

Topic 2.1: Polarity of Water

Water is a highly polar covalent compound. The Oxygen has a partial negative charge, and the Hydrogens/hydrogen end have a partial positive. This makes it better as a dissociative solvent for ionic compounds. Dissociation is when ionic compounds separate into their respective ions with the same charges they have while in the compound. These individual atoms (like Na^+ or Br^- from $NaBr$) and their parent compounds (such as $NaBr$) are called electrolytes.

Dissociation occurs in significant amounts in aqueous solutions because of a process called hydration. Water molecules' partially negative oxygen atoms surround positive ions (cations) like Na^+ to help them stay in solution and prevent reassociation. In order to cancel the electric charge, more than one water molecules surround the ions. Similarly, the partially negative hydrogen side surrounds negative ions (anions) like Cl^- .

Topic 2.2: Ionization

Water can also dissolve some covalent compounds by ionizing them—a chemical reaction. For example, sulfur dioxide (SO_2 , a covalent compound) takes OH^- from the water molecule to make itself a negative ion. This also occurs with ammonia and hydrochloric acid. Note that this, like dissociation, only occurs to water-soluble molecules.

Topic 2.3: State of Matter in Chemical Reactions

As described in the lesson on chemical equation symbols, symbols like (*g*) and (*l*) describe the state of matter of reactants and products (gas and liquid, respectively), and (*aq*) describes reactants or products dissolved in an aqueous solution. Where insoluble products form (remaining as (*s*) instead of becoming (*aq*) like with $AgCl$), gases either bubble out of the solution and solids form precipitates separate from but within the solution.

Topic 2.4: Double-Displacement Reactions

Double-displacement reactions occur when ions switch places in ionic compounds. These often occur in aqueous solutions, and when they do, ionic compounds in aqueous solution can be represented as their individual ions in aqueous solution (e.g. $NaCl(aq) = Na^+(aq) + Cl^-(aq)$ but $AgCl(s) \neq Ag^+(s) + Cl^-(s)$).

After these dissociated ions are rewritten, any ion which is on both sides in the same phase can be eliminated (e.g. $H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^- + H_2O(l)$ becomes $H^+(aq) + OH^- \rightarrow H_2O(l)$). This is called a net ionic equation. Notably, if all reactants and products are aqueous in a double-displacement reaction, the net ionic equation is empty and thus no real reaction has occurred. This leads us to show that double-displacement reactions in solution occur only when a gas, precipitate, or pure liquid are formed.

Topic 2.5: Types of Electrolytes

Strong electrolytes and weak electrolytes are the two types of electrolytes. Strong electrolytes dissociate or ionize completely like $HCl(g)$ or $KOH(s)$. This means that they conduct electricity very well when in solution. Weak electrolytes, on the other hand partially dissociate or ionize because their reactions are easily reversible—meaning that a significant amount of reactant is left in the solution. Nonelectrolytes aren't really electrolytes because they don't dissociate or ionize in solution. These solutions do not conduct electricity.

Lesson 3: Solutions and Solubility

Topic 3.1: The Dissolving Process

Dissolution does not occur instantaneously. It's a process which requires that solute molecules or ions are attracted to solvent molecules, causing the solvent to try and surround the solute. This pulls apart the solute from the other solute molecules, and completes the mixing process. The necessary attraction between solute and solvent is why the “like dissolves like” rule tends to accurately describe dissolution of solutions (e.g. water, a highly polar covalent molecule easily dissolves highly polar ionic compounds).

Topic 3.2: The Rate of Dissolution

The rate of dissolution is how quickly a solid solute dissolves in a liquid solvent. This increases proportionally with stirring, surface area, and temperature.

Topic 3.3: Solubility

Solubility is the “amount of solute which will dissolve in a volume of solvent at a given temperature and pressure.” This changes only with the solute's and solvent's identity. This is different from the dissolution rate and follows the like dissolves like rule significantly.

The concentration of a solute compared to the maximum concentration based on solubility defines three types of solutions: unsaturated solutions have less solute than the “maximum,” saturated solutions have exactly an equal amount, and supersaturated solutions have higher than the maximum. Supersaturated solutions are highly unstable and are typically created by slowly cooling a saturated solution at a higher temperature.

Topic 3.4: Environmental Conditions' Effects on Solubility

Increasing temperature increases the solubility of solids and liquids but decreases the solubility of gases. Increasing pressure, however, increases the solubility of gases and does not affect the solubility of solids and liquids.

Solubility graphs show the change in solubility with temperature. These are determined experimentally because they are different for different solutes.

These can be used to determine if a given solution is saturated, unsaturated, or supersaturated at a specific temperature.