Chemistry

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Contents

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 $1.66 * 10^{-24}g$.

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

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 $N = A - Z$

$$^{27}_{13}\mathrm{Al}$$

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 $35(75.78\%) + 37(24.22) \simeq 35.45amu$

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

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: $^{235}_{92}U +^{1}_{0} n \rightarrow^{94}_{38} Sr +^{140}_{54} Xe + 2^{1}_{0}n$

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 $^{2}H +^{3} H \rightarrow^{5} He \rightarrow 17.6MeV +^{4} He + n$. Another common pathway is the Carbon-Oxygen-Nitrogen cycle, which uses those atoms to turn 4 independent protons into a ^{4}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

$$\underset{2s}{\underline{\upharpoonleft\downharpoonright}} \quad \underset{2p}{\underline{\upharpoonleft}\ \underline{\upharpoonleft}\ \underline{\ \ }}$$

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 $1s^2 2s^2$

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 occuring 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 13 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 $4s^23d^104p^3$) 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 bcomes 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 eachother 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 partilces
- All collisions between particles are perfectly elastic (no energy is lost)
- Average kinetic energy is only dependent on temperature.

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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. $\frac{P_1V_1}{N_1T_1} = \frac{P_2V_2}{N_2T_2}$.

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 pairs 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 pproperties (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.

Lesson 6: Ionic Bonding

Topic 6.1: 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 6.2: 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 6.3: 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 bound of $A_B B_A$ (simplified to a ratio, and assuming A and B represent the respective elements).

Topic 6.4: 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 6.5: 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 7: Covalent Bonding

Topic 7.1: 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 visual electrons with eachother.

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 7.2: 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 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 7.3: 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 7.4: 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 7.5: 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 7.6: 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 7.7: 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 8: Metallic Bonding

Topic 8.1: 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 8.2: 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 8.3: 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 8.4: 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).