Honors Government

Hashem A. Damrah

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Contents

CHAPTER ONE

Matter

Unit 1

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Lesson 1: The Science of Chemistry **Unit 1**

Knowledge that's considered science

Definition 1 (Chemistry)**.** Chemistry is the study of composition and the structure of materials and the changes they undergo.

This includes studying materials found in our:

- Ocean
- Atmosphere
- Environment
- Underground
- Etc ...

Chemistry is one of many fields of science that investigate phenomena such as climate change to discover their causes.

Definition 2 (Science)**.** Science must be based on empirical observations, experimentation, explanations based on logical reasoning.

- OBSERVABLE: Science attempts to explain natural phenomena by analyzing and observing the world and testing ideas about it.
- TESTABLE: Science must be able to answer a testable question using observation and experimentation. Investigations must produce empirical evidence that can be observed or measured to be considered science.

- REPLICABLE: Empirical evidence can be replicated, or reproduced, and verified by other scientists if they conduct the same tests and under the same conditions.
- RELIABLE: The more an experiment is repeated, with the same outcomes, the more reliable the evidence becomes. Evidence with bias also increases its reliability.
- FLEXIBLE: Science is an ever-changing body of knowledge as new observations are made through experimentation. As new information is discovered, new evidence can add to current evidence, allowing scientists to improve their theories.

Answering questions with Science

Many fields of knowledge, such as philosophy and art add to our view of the world. They help us appreciate the beauty of the world, interact with one another and decide what is wrong and right. But, since there aren't any observations and tests being applied in those fields and everything in them are just beliefs and opinions, then it cannot be called science.

These two questions will help you determine if the question can be answered with science:

- If the question is asking about an opinion or a moral value, it's not something that can be measured using scientific process. Therefor, it cannot be answered with science.
- If the answer to the question cannot be tested and observed, it is not considered science.

A scientific method is a series of steps for investigating questions and testing ideas. There are several versions of the **scientific method**, but all of the versions are based on rational thinking, inquiry, and experimentation. The **scientific method** includes five main steps:

- Question: A scientific method always starts with a question. As long as it's testable, you can use science to find the answer.
- Research: It's important that you always explore what other scientists before have observed and discovered because that information may be able to solve your current question.
- Hypothesis: Most hypotheses are made with the **"if/then"** statements. If "this" happens, then "that" will take place. This helps you know how one thing can affect another and gives you variables to test.
- Testing: An experiment allows you to test your hypothesis to determine if you're correct or incorrect.
- Independent Variable: This variable is the factor the scientist has chosen to change in the experiment. A good experiment should change only one independent variable because this allows the scientist to focus only on the outcome of that specific variable.

- Dependent Variable: This variable is the factor that changes in response to the independent variable in an experiment.
- Controlled Variables: These variables are the factors the scientist chooses to keep constant over the course of the experiment. By keeping all other factors constant, anything happens to the dependent variable is caused by the independent variable.

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Lesson 2: Measuring Matter **Unit 1**

Measurements and Units

All measurements are made of **2** parts:

- A number.
- A unit.

Definition 3 (Matter)**.** Anything that takes up space, whether big or small, is **Matter**.

Definition 4 (Volume)**.** In science, the amount of space an object occupies is called **volume**

Definition 5 (Mass)**.** **Mass** is a measure of the amount of matter in an object. This means, the more **Matter** an object has, the object's **Mass** increases.
Scientists measure the **Mass** of an object using the base **Unit Gram**.

Systems of Measurements

The **U.S. Customary System (English System)** consists of measurements we use every day to describe:

- How much we have.
- What size we want.
- How far we want to go.
- Etc ...

The **United States** is the one of three countries in the **whole world** that use this system.

Type	U.S. Customary System	Metric System
Length	inch, foot, mile	meter
Mass/Weight	ounce, pound, ton	gram
Volume	pint, quart, gallon	liter
Temperature	degrees Fahrenheit	degrees Celsius

Table 1.1: The **U.S. Customary System (English System)**

The **Metric System** is a system of measuring units based on the power of 10. It's the preferred system used in science measurement because so many countries already use metrics as their standard measurement system.

One reason for the widespread use of the **Metric System** is that it seems to simplify the number of measurements. Instead of multiple units for the same measurements, such as:

- Inches.
- Feet.
- Miles.

for length, the **Metric System** uses only one word for the length, the **Meter**.

For larger or smaller measurements, metric prefixes are used rather than changing the unit of measure or using fractions, like quarter of an inch. A prefix works with any unit of measure in the **Metric System**:

- Meter.
- Gram.
- Liter.

Prefix	Symbol	Multiplier	Example
Kilo-	k	$1,000$	Kilometer (km)
Hecto-	h	100	Hectometer (hm)
Deca-	da	10	Decameter (dam)
Unit		1	Meter (m)
Deci-	d	0.1	Decimeter (dm)
Centi-	c	0.01	Centimeter (cm)
Milli-	m	0.001	Millimeter (mm)

Table 1.2: The **Metric System Prefix References**

The **International System of Measurement of Units** is the primary system used by scientists and engineers. It will also be used in **Chemistry**. In this system, scientists have specified what units should be used as the seven base units of measurement from which all other units are derived.

Quantity (Symbol)	Name of Unit (Symbol)
Length (l)	Meter (m)
Mass (m)	Kilogram (kg)
Time (t)	Second (s)
Electric Current (i)	Ampere (A)
Thermodynamic Temperature (T)	Kelvin (k)
Amount of Substance (n)	Mole (mol)
Luminous Intensity (l_v)	Candela (cd)

Table 1.3: The **Metric System Prefix References**

Converte between or within Measurement Systems

Definition 6 (Dimensional Analysis)**.** Analyzing the relationships between different quantities, identify their units of measurement, and convert their quantities for equal comparison.

Dimensional Analysis uses a ratio, or a conversion factor to change units of measurement while maintaining the value of measurement.

In a **Conversion Factor**, the **numerator** and **denominator** equal the same amount, just in different units.

For example, to convert 3 miles into feet, you would multiply 3 miles by a conversion factor created from the units of miles and feet. Because 1 mile equals $5,280$ feet, two conversion factors can be created

$$1 = \frac{1 \text{ mile}}{5,280 \text{ feet}} \text{ or } 1 = \frac{5,280 \text{ feet}}{1 \text{ mile}} \tag{1.1}$$

Multiplying any number, such as 3 miles, by either of these fractions does not change the number's overall value.

Using the conversion factor for converting derived units

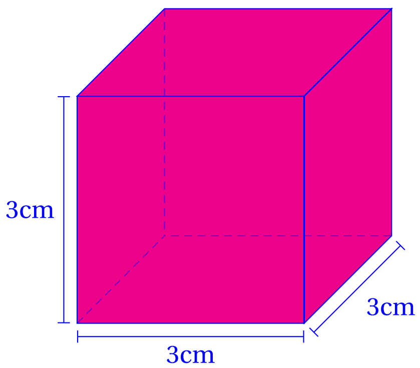

Figure 1.1: Cube that's $3 \times 3 \times 3$

Definition 7 (Derived Units)**.** Some measurements you are familiar with, such as speed, contain more than one **Unit**.

A very common **Derived Unit** is **Volume**. For liquids and gasses, **Volume** tends to be measured in liters. But, for regular-shaped solids, measurements and calculations can determine their **Volume**. To calculate the **Volume** of the cube in Figure 1.1, just raise the side measurement to the third power.

$$\begin{aligned} V &= (3cm)^3 & (1.2)\\ &= (3 \times 3 \times 3)cm & (1.3)\\ &= 27cm^3 & (1.4) \end{aligned}$$

Now, let's learn how to convert **Derived Units**:

Length	Weight	Liquid Capacity
1 foot = 12 inches	1 pound = 16 ounces	1 cup = 8 fluid ounces
1 yard = 3 feet	1 ton = 2,000 pounds	1 pint = 2 cups
1 mile = 5,280 feet		1 quart = 2 pints
1 mile = 1,760 yards		1 gallon = 4 quarts

Table 1.4: U.S. Customary Units of Measure

Length	Weight/Mass	Liquid Capacity
1 foot = 2.54 centimeters	1 pound = 0.454 kilograms	1 gallon = 3.785 liters
1 meter = 39.37 inches	1 ton = 2.2 pounds	1 liter = 0.264 gallons
1 mile = 1.609 kilometers	1 ounce = 28.35 grams	1 liter = 1,000 cubic centimeters
1 kilometer = 0.6214 miles		1 milliliter = 1 cubic centimeter

Table 1.5: Metric Units of Measure

Prefix	Symbol	Multiplier	Example
Kilo-	k	$1,000$	Kilometer (km)
Hecto-	h	100	Hectometer (hm)
Deca-	da	10	Decameter (dam)
Unit		1	Meter (m)
Deci-	d	0.1	Decimeter (dm)
Centi-	c	0.01	Centimeter (cm)
Milli-	m	0.001	Millimeter (mm)

Table 1.6: The **Metric System Prefix References**

Example (Fractional Unit)**.** If you start with a number that has a factional unit, there may be more steps to the **Dimensional Analysis**. It doesn't have to be harder, you still want to make sure that you cancel out all the units except the units you want in your final answer. Look at the following problem to see how to change the numerator and/or denominator of a fractional unit.

In this problem, you've been asked to convert from the fractional unit kilometers per hour to the unit centimeters per second. This means that you'll need to convert the numerator and denominator to get the final answer.

When you have a given that has a fractional unit, it can be helpful to write the measurement as a fraction with an one in the denominator:

$$\frac{0.36km}{1h} \tag{1.5}$$

This makes it more obvious that there are two numbers that need to convert: $0.36km$ and $1h$.

Once you know what needs to be converted, you can determine those relationships that will help you go from kilometers to centimeters and hours to seconds.

Since you need to change both the numerator and the denominator, you can start with the numerator first. The one hour is still in the denominator, but you can ignore it until you get kilometers converted to centimeters:

$$\frac{0.36\,\cancel{\text{km}}}{1h} \times \frac{1,000\,\cancel{\text{m}}}{1\,\cancel{\text{km}}} \times \frac{100cm}{1\,\cancel{\text{m}}} \tag{1.6}$$

Notice that the units that need to be canceled (km and m) are all diagonal from each other. That leaves centimeters as the only unit that does not cancel. It is in the numerator, where it needs to be for the final answer.

Now that we've changed the numerator from the unit km to cm on top, you can continue the problem by changing the denominator from hr to sec:

$$\frac{36,000cm}{1\,\cancel{\text{h}}} \times \frac{1\,\cancel{\text{h}}}{60\,\cancel{\text{min}}} \times \frac{1\,\cancel{\text{min}}}{60sec} \tag{1.7}$$

Notice that you need to put hours on the top of the next conversion in order to cancel out with ours in the denominator from hr to sec

Notice that you need to put hours on the top of the next conversion in order to cancel out with hours in the denominator of the given. In the next step, the minutes unit is canceled to leave seconds in the denominator – where it needs to be for the final answer:

$$\begin{aligned} \frac{36,000cm}{1\,\cancel{\text{h}}} \times \frac{1\,\cancel{\text{h}}}{60\,\cancel{\text{min}}} \times \frac{1\,\cancel{\text{min}}}{60sec} &= 10\frac{cm}{s} && (1.8)\\ &= 0.36\frac{km}{h} && (1.9)\\ &= 10\frac{cm}{s} && (1.10) \end{aligned}$$

Example (Powered Unit)**.** Some units are raised to a power. **Area** and **Volume** of solids are the common ones.

Area is measured in squared length, such as the square meter, m^2, while **Volume** is measured in cubic length, such as the cubic meter, m^3. Because one cubic centimeter equals 1 milliliter, that conversion is quite simple. However, there are times you must convert from one cubic meter to a cubic centimeter. Let's look at an example:

Convert the **Volume** $5.0m^3$ to the unit cm^3.

Just like any unit conversion, we look at the starting and ending units, then think about what we know that can help us convert the units. You haven't learned an equivalent relationship between cubic metres and cubic centimeters, but you know the relationship between the meters and centimeters:

$$1m = 100cm \tag{1.11}$$

We simply use the relationship we already know and cube the entire thing. This makes the fraction a three dimensional **Volume** measurement instead of a one dimensional length measurement:

$$\left(\frac{100cm}{1m}\right)^3 \tag{1.12}$$

After we do that, we can setup the problem, by starting with the given measurement and adding the relationship we already know.

The units m^3 and m cannot cancel out because they are not exactly the same. The entire conversion factor must be cubed in order to make the units match:

$$5.0m^3 \times \frac{100cm}{1m} = 5.0m^3 \times \left(\frac{100cm}{1m}\right)^3 \tag{1.13}$$

$$= 5.0m^3 \times \left(\frac{1,000,000cm^3}{1m^3}\right) \tag{1.14}$$

Now that the cubic meter units are able to cancel, it's time to calculate the answer. Because the entire fraction is being cubed, the numbers and units must all be cubed:

$$5.0m^3 \times \frac{1,000,000cm^3}{1m^3} \times \frac{5.0 \times 1,000,000cm^3}{1} = 5,000,000cm^3 \text{ or } 5 \times 10^6 cm^3 \tag{1.15}$$

Therefore, $5.0m^3$ is equivalent to $5 \times 10^6 cm^3$.

Sep 10 2021 Fri (06:53:22)

Lesson 3: Energy and Temperature **Unit 1**

Different Forms of Energy

Particles in motion can cause change due to their motion. For example:

Definition 8 (Energy)**. Energy** is the ability to cause change or the ability to do work. **Energy** is measured in the **SI Unit Joules**:

$$\text{Joule} = \frac{kg \times m^2}{sr} \tag{1.16}$$

Definition 9 (Work)**. Work** is the process of causing matter to move against an opposing force. When matter moves, it's position is changing. For that to happen, it requires **Energy**.

All forms of energy can be broadly classified as either:

- **Potential Energy**.
- **Kinetic Energy**.

Definition 10 (Potential Energy)**. Potential Energy** is the energy an object has because of it's current position or composition. For example, you have someone standing on a plane $25,000ft$ above ground before he sky dives which gives in **Potential Energy**.

Potential Energy is also known as **Stored Energy** because it's energy that is there, but hasn't been used. You can think about it as, winding up a toy car, then once you let it go, all of the energy comes out as **Kinetic Energy**.

Definition 11 (Kinetic Energy)**. Kinetic Energy** is energy an object has due to its motion. Like I mentioned before, when you wind up a toy car and then let it go, that's an example of **Kinetic Energy**.

Definition 12 (Mechanical Energy)**. Mechanical Energy** is the movement of objects from one place to another

EXAMPLES:

- Turning wheels.
- Wind blowing.

Definition 13 (Stored Mechanical Energy)**. Stored Mechanical Energy** is the energy stored in objects by application of force.

EXAMPLES:

- Compressed spring.
- Pulled back string of a bow before shooting the arrow.

Definition 14 (Electrical Energy)**. Electrical Energy** is the movement of electrical charge.

EXAMPLES:

- Electricity running though wires.
- Lighting.

Definition 15 (Electrical Potential Energy)**. Electrical Potential Energy** is the stored energy of electric charges due to their position and interaction around other charges.

EXAMPLES:

- Non-moving charges in circuits.

Definition 16 (Radiant Energy)**. Radiant Energy** is electromagnetic energy that travels in waves.

EXAMPLES:

- X-Rays.
- Microwaves.
- Visible light.
- Radio waves.

Definition 17 (Chemical Energy)**. Chemical Energy** is stored in the bonds of atoms and molecules.

EXAMPLES:

- Biomass.

- Natural gas.
- Propane.

Definition 18 (Sound Energy). **Sound Energy** is the movement of energy through substances in longitudinal waves. Sound is produced when a force causes an object to move or vibrate, which is when energy is transferred through the object in a wave.

Definition 19 (Nuclear Energy). **Nuclear Energy** is energy stored in the nucleus of an atom. This is the energy that holds atoms together and can be released in processes called fusion and fission.

Definition 20 (Thermal Energy). **Thermal Energy** also known as **heat**, **Thermal Energy** is the internal energy in a substance caused by vibration of atoms and molecules. **Geothermal Energy** is an example of **Thermal Energy**. Heat is also a byproduct of a lot of energy conversions.

Definition 21 (Gravitational Energy). **Gravitational Energy** is the **Potential Energy** of position. The rock resting at the top of a hill has **Gravitational Potential Energy**.

Although there are many forms of energy, the **Law of Conservation of Energy** states that ***"Energy can be converted from one form to another, but it's not created or destroyed in ordinary physical and chemical processes."***

Energy being conserved during Physical and Chemical Change

Energy rarely stays in its current form. Instead, it is converted to different forms through **physical** and **chemical** processes. As it changes from one form to another within these processes, more energy is not created, and none of the energy is destroyed; it simply changes form.

Heat Vs. Temperature

To understand the relationship between heat and the motion of particles, let's look at how thermal energy affects the interaction between and the motion of sugar and coffee particles.

For the sugar crystals to dissolve, the crystals need to break apart into smaller molecules. When we stir, we help break down the solid particles by causing them to move and change faster. When heat transfers to the molecules in the coffee, they vibrate faster, increasing their kinetic energy. This increases the number of collisions between the particles to break down sugar molecules.

The average kinetic energy of molecules is what determines the temperature of a substance. So, hotter objects have faster-moving particles and higher temperatures. Temperature measures the averages because moving particles can absorb and release heat with each collision as heat is converted to kinetic energy and back.

Sep 13 2021 Mon (10:54:11)

Lesson 4: Properties of Matter **Unit 1**

Differences between each State of Matter

There are four known states of matter:

- Solid
- Liquid
- Gas
- Plasma

> **Definition 22** (Phase)**.** Each one of these stages is known as a **Phase**. Each one has it's own properties, such as:
>
> - Shape
> - Particle Motion

Solid substances can be **hard**, or **soft**, but they have low **compressibility** and usually maintain their shape on their own.

On the other hand, a substance in a **liquid** will take the shape of the container it's poured into and fill the container from the **bottom up**.

A **gas substance** will fill up the entire container spreading out evenly because their particles can be forced closer together with pressure.

Like **gas**, **plasma** can spread evenly in a container and be compressed, but its composition differs. **Plasma Particles** are positively and negatively charged, which allows plasma to conduct electricity and create magnetic fields.

A **physical property** can be changed without altering the identity of a material. Each physical properties can be classified as either an **Extensive Property**, or an **Intensive Property**.

- Shape
- Volume
- Length
- Mass

are classified as **Extensive Physical Properties**. An extensive property is a physical property that depends on the sample size.

Describing matter using Physical Properties

Example (Extensive Physical Properties)**.** For example, if you have a large amount of metal and a small amount of the same metal, these two samples will have different volumes.

- Magnetism
- Density
- Melting
- Boiling
- Points
- Color

are all classified as **Intensive Physical Properties**. Intensive properties do not depend on the size of the samples. No matter how much or how little of the substance you have, an intensive physical property stays the same.

Example (Intensive Property). For example, if you have a large amount of metal and a small amount of the same metal, these two samples will have the same density.

Definition 23 (Phase State). The phase:

- Solid
- Liquid
- Gas
- Plasma

of a substance is a physical property.

Definition 24 (Density). The ratio of mass per volume is a physical property that does not depend on the size of the sample.

Definition 25 (Ductility). The ability to be pulled or stretched to make wire is a physical property of metals.

Definition 26 (Malleability). Some substances have the ability to be:

- shaped
- Dented
- Extended

by beating with a:

- Hammer
- Rolling

Definition 27 (Melting Point)**.** The temperature at which a solid becomes a liquid is its **Melting Point**.

Definition 28 (Solubility)**.** A substance is **Soluble** if it's able to dissolve in another substance.

Definition 29 (Boiling Point)**.** The temperature at which a liquid becomes a gas is its **Boiling Point**.

Definition 30 (Compressibility)**.** **Compressibility** is a physical property. It describes how much the volume of matter decreases under pressure. Substances with high compressibility can be squeezed or flattened to fit into smaller spaces.

Definition 31 (Electrical and Thermal Conductivity)**.** The ability to transmit electricity or heat through the structure of a substance is a physical property shared by all metals.

Calculating Density

Because density is an intensive physical property, its value for a substance does not change based on the sample size of that substance. Because of this, density is often used to identify unknown materials. Density measures the degree to which something is compacted, or the quantity of mass per unit volume.

$$density = \frac{mass}{volume} \tag{1.17}$$

Example (Example One)**.** The sample of a cork has a mass of $2.88g$ and a volume of $12.0mL$. What's the density of the cork?

We have the following information:

$$mass = 2.88g \tag{1.18}$$

$$volume = 12.0mL \tag{1.19}$$

We are asked to calculate density. So, use the formula:

$$density = \frac{mass}{volume}$$

Substitute the numbers and divide:

$$d = \frac{2.88g}{12.0mL} \tag{1.20}$$
$$= 0.25\frac{g}{mL} \tag{1.21}$$

Example (Example Two)**.** This sample of marble block is in the shape of a **cube**. It has a mass of $22.4kg$ and a side length of $2m$. What's the density of the block?

First, find the volume of the cube:

$$V = (2m)^3 \tag{1.22}$$
$$= 8m^3 \tag{1.23}$$

The given information:

$$\text{mass} = 22.4kg \tag{1.24}$$
$$\text{volume} = 8m^3 \tag{1.25}$$

So, use the formula:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{1.26}$$

Substitute the numbers and divide:

$$d = \frac{22.4kg}{8m^3} \tag{1.27}$$
$$= 2.8\frac{kg}{m^3} \tag{1.28}$$

Definition 32 (Reactivity)**. Reactivity** is a chemical property. It refers to the readiness of a substance to undergo a chemical change.

There are some substances, like gold or platinum, that do not react so easily with other substances. The low reactivity of these elements is perfect for jewelry making because they will not rust or tarnish when exposed to other substances.

An element with high reactivity, like fluorine - a poisonous gas - can spontaneously explode when exposed to other substances.

Definition 33 (Flammability)**. Flammability** is a chemical property. It describes how easily a substance can be set on fire. This is very important when it comes to building safety.

For instance, some Styrofoam-based insulation products have higher flammability than other fire-resistant products. Builders avoid using flammable materials in homes.

Alcohol and gasoline are two liquids that are highly flammable.

Definition 34 (Toxicity)**. Toxicity** is a chemical property. It describes the ability of a substance to damage or harm an organism.

Higher toxicity equals more harm. For example, exposure to a gram of mercury, an element that is liquid metal, can cause death to humans.

There are some creatures that are labeled venomous or poisonous because of their ability to harm other living creatures due to the toxic substances they naturally produce.

Definition 35 (Heat of Combustion)**.** The **heat of combustion** is a chemical property. It measures how much heat is given off when a substance is burned.

Definition 36 (Corrosion)**. Corrosion** is the irreversible damage or destruction of a material due to a chemical or electrochemical reaction.

Definition 37 (Decomposition)**.** Some compounds decompose into more than one different element or compound as bonds are broken. Hydrogen peroxide (H_2O_2), a chemical you may have around your house, is stored in a dark container to slow down **decomposition**, which can be caused by heat and light.

Sep 13 2021 Mon (14:32:49)

Lesson 5: Changes of Matter **Unit 1**

Differences between Physical and Chemical changes

Definition 38 (Physical Change)**.** A **Physical Change** is a change in a substance from one form to another without changing its composition. A physical change can be a change to any physical property of an object, such as its size or shape.

Definition 39 (Chemical Change)**.** A **Chemical Change** is change in a chemical substance into one more different substance. This type of change occurs at the molecular level, where atoms are re-arranged to form new molecules during chemical reactions. Unlike **Physical Change**, **Chemical Change** is permanent.

All mater has physical and chemical properties. Understanding these how properties are affected when matter changes will allow you to know if a substance has undergone a physical or chemical change.

Here are some clues that will help you tell if something has undergone a chemical change:

1. production of flame
2. color change
3. bubbling or fizzing
4. temperature change
5. smoke
6. production of light
7. formation of a substance in a different state

Sep 21 2021 Tue (09:21:04)

Lesson 6: Pure Substances and Mixtures **Unit 1**

Classifing matter as a Mixture or Pure Substance based on its Properties

All matter can be classified into two main categories

- Mixtures
- Pure Substances

Definition 40 (Matter)**. Matter** is anything that takes up **Mass** and **Space**.

Mixtures

A mixture is created when solids, liquids, or gases mix with one another. Mixtures can occur with substances that are in the same states of matter or of different states of matter. Each substance in a mixture maintains its individual composition and property.

Definition 41 (Homogeneous)**.** In this type of mixture, just one substance looks the same throughout. Dissolving sugar in a glass of water forms a **homogeneous** mixture, as the sugar seems to disappear. Even when examined under a microscope, the sugar water appears the same throughout.

Some examples of **Homogeneous Mixtures** are:

- Air
- Perfume
- Bleach
- Steel

Definition 42 (Heterogeneous)**.** Candies in a mixture will not be dispersed evenly. Heavier pieces may sink to the bottom of the pile, and smaller pieces may be at the edges. Vegetable soup is an example of a heterogeneous mixture because each spoonful will have a slightly different mixture of vegetable pieces and seasoning.

Some examples of **Heterogeneous Mixtures** are:

- Rocks
- Cereal
- Sand

Definition 43 (Physical Separation)**.** Mixtures contain compounds and elements that can be physically separated from one another. For example, you can pull out your favorite fruit from a fruit salad and leave the others in the bowl. Or you can boil a saltwater mixture to cause the water to evaporate and leave salt particles.

Pure Substances

A pure substance consists of a single element or type of compound. For example, gold is a **Pure Substance** because it's made from a single element: gold.

Definition 44 (Elements)**.** An **Element** is a substance that cannot be broken down into simpler substances. All elements on the periodic table of elements are examples of **Pure Substances**.

Definition 45 (Compounds)**.** A **Compound** is matter composed of two or more **Elements**. A molecule is the smallest unit of a **Compound**.

Some examples of **Compounds** are:

- Alcohol
- Water
- Baking Soda
- Glucose

Definition 46 (Chemical Separation)**.** Pure substances made from compounds can only be chemically separated into the elements that make them up. The compound table salt ($NaCl$) is made from the elements **Sodium** and **Chlorine**. A chemical reaction would be needed to separate these two elements from one another.

Sep 21 2021 Tue (10:37:32)

Lesson 7: Laboratory Techniques **Unit 1**

Different kinds of measurement tools

Definition 47 (Test tubes and test tube holder)**.** Used to hold small amounts of a substance

Definition 48 (Dropper pipettes)**.** Used to transport a measured volume of liquid

Definition 49 (Ring clamp and stand)**.** A metal stand with a clamp used to support a container of material for testing

Definition 50 (Thermometer)**.** An instrument used for measuring temperature

Definition 51 (Florence flask)**.** A container used for uniform heating and boiling of liquids

Definition 52 (Beaker with glass stirring rod)**.** A cylindrical glass container for holding substances and a glass rod for safe stirring of materials

Definition 53 (Erlenmeyer flask)**.** A flat-bottomed flask for substances

Definition 54 (Balance scale)**.** A device that compares the weight of an unknown object on one scale pan to the weight of a standard mass on the other scale pan

Definition 55 (Bunsen Burner)**.** A heat source for heating substances in the laboratory

Definition 56 (Mortar and pestle)**.** Used to crush ingredients or substances by grinding them into a powder

Definition 57 (Graduated cylinder)**.** Equipment used to measure the volume of a liquid and some solids using water displacement

Definition 58 (Funnel)**.** Used for guiding liquid or powder into a small opening

Definition 59 (Crucible)**.** A ceramic or metal container used for melting or heating substances

Sep 22 2021 Wen (06:30:13)

Lesson 8: Honors Scientific Knowledge **Unit 1**

Difference between Science and Pseudoscience

Science is often considered the pursuit and acquisition of knowledge. Therefore, any actions that mimic this search for answers can look like science. However, the key difference between an explanation built with science and one built without it is the gaining of evidence through experimentation and the scientific method.

Sometimes a practice or belief claims to be science, but it does not follow the scientific method or is not proven reliable through experimentation. This is called pseudoscience. Let's learn why these examples of pseudoscience are not considered examples of science.

Scientists seek to understand the world around us and to learn how things work. To find answers, they investigate all sorts of natural phenomena and conduct experiments to collect data for analysis and interpretation. But to prove that their interpretations are valid, and in turn become scientific explanations of natural phenomena, their investigations must meet certain criteria.

- Following Logic
  Using logic to interpret experimental data is important to the validity of a scientific explanation. For example, if data show that water evaporates with the addition of heat, it would not be proper logic to assume the water spontaneously evaporated.

- Peer Review
  Frequent examinations by scientists result in some ideas being refuted and replaced with other ideas. This frequent examination and testing makes scientific explanations more valid and durable over time.

- Global Access
  Scientific knowledge is constantly changing and developing because new observations or predictions are made every day. All scientific knowledge should be examined and re-examined using the steps of the scientific method to collect new empirical evidence.

- Rules of Evidence
  It is important that scientists share their investigations and conclusions with others so they can be tested and used by scientists all over the world.

Hidden Phenomena

A scientific explanation should be observed, tested, or measured to be considered valid and reliable. Yet, there are some natural events that just can't be observed or tested using the technology available. For instance, in the 16th century Nicolaus Copernicus first proposed a mathematical heliocentric model. It theorized that the sun, not Earth, was the center of our solar system. The model was supported by empirical evidence collected from astronomers tracking objects in the night sky. But the model could not be completely confirmed until we sent probes and people into space centuries later.

Also, there are things in ancient history, before the written word, that scientists can only theorize about based on studies of items from that period. For example, the existence of dinosaurs is supported by numerous sources of their skeletal remains collected all around the world. Although no one has seen a living dinosaur, studies of these remains support that they once lived in vast numbers.

Theories and Hypothesis

Within the body of scientific knowledge are two very important components: hypotheses and theories. Both types of scientific knowledge are developed using background research and can change over time using the scientific method. The difference is that scientific theories are well-tested hypotheses. They describe why things happen using empirical evidence collected and tested by multiple scientists using multiple kinds of investigations over a significant period. Hypotheses, on the other hand, are predictions or tentative explanations of phenomena that have not yet been fully tested or have been dis-proven.

CHAPTER TWO

Atoms and Elements

Unit 2

Oct 06 2021 Wen (07:20:33)

Lesson 1: Atomic Theory **Unit 2**

In the $1800s$, gathering evidence for an atomic model was challenging. At that time tools didn't exist for scientists to get a good hard look at atoms. But that didn't stop **John Dalton** from observing atomic behavior. He devised a way to indirectly observe atomic nature using things he could measure and observe, such as

- Temperatures
- Atmospheric conditions
- Weather patterns

Here are some of the things that he found:

Definition 60 (All matter is composed of extremely small particles called atoms.)**.** The word *"atom"* comes from a Greek word meaning *"cannot be cut into smaller pieces"*. Atoms have their own internal structure that can vary in mass and composition. The types of atoms that make up matter determine its properties.

Definition 61 (Atoms of a given element are identical in size, mass, and other properties.)**.** An element is a substance that cannot be broken down into simpler substances. Atoms of a given element are identical. For example, all silver atoms are exactly the same. A silver atom may have some similarities with a gold atom, but there will always be some differences, such as mass or other properties.

Definition 62 (Atoms cannot be subdivided, created, or destroyed in chemical reactions.)**.** According to the law of conservation of mass, matter cannot be created or destroyed. However, scientists will discover that atoms can be subdivided under unique circumstances.

Definition 63 (Atoms of different elements can combine in simple, whole number ratios to form chemical compounds.)**.** John Dalton's observations eventually led to the law of multiple proportions, which explains that atoms combine to form compounds in simple, whole-number ratios. For example, water is made of two atoms of hydrogen and one atom of oxygen. There's no such thing as a half atom within a compound.

Definition 64 (In chemical reactions, atoms are combined, separated, or rearranged.)**.** The bonds between atoms are broken, rearranged, and reformed into new compounds during chemical interactions.

There's only so much you can study about atoms from a macroscopic level. To understand the atomic nature of matter, scientists had to get microscopic. Let's take a closer look at the important atomic discoveries made by J. J. Thomson, Ernest Rutherford, Niels Bohr, and James Chadwick—no microscopes necessary.

1897-1906 – Discovery of Electrons

Experiment J. J. Thomson loved working with cathode rays—the glowing beams of light we once used in television and computer screens. In 1897, after many cathode ray tests, J. J. Thomson discovered smaller, negatively charged particles inside the atom. He called them corpuscles because they were so small, but we now call them electrons.

Conclusion Thomson hypothesized that these corpuscles were scattered within a positively charged atom, just like plums are mixed throughout plum pudding—hence the name of his model. (You may think of it as chocolate chips scattered in cookie batter.)

1908-1913 – Rutherford and the Nucleus

Ernest Rutherford was a student of J. J. Thomson, so he supported the plum pudding model. But he soon discovered that the model didn't quite fit.

Experiment In Rutherford's experiments, he made observations about the behavior of positively charged particles as they were radiated onto a piece of gold foil. He noticed some of the particles did not pass straight through the foil as he expected. Some particles scattered, and some bounced back.

Conclusion Rutherford proposed that there must be something big and positively charged in the center of each gold atom to cause the radiation particles to do this. He named this center the nucleus of the atom. He also proposed that the nucleus was composed of two particles, one that was positively charged (proton) and one that was neutral.

1913 – Bohr's Orbitals

Experiment Niels Bohr worked with Ernest Rutherford on atomic structure. After observing the spectrum of colors that were emitted when a gas was excited (electrons gained energy), he believed the different wavelengths of color represented the energy levels an electron could exist within.

Conclusion Bohr proposed that electrons moved in specific orbits around a positive nucleus based on their energy level. He predicted there were many energy levels in which electrons could exist. They are not randomly scattered, as scientists first predicted.

1927-1932 – Chadwick and the Neutrons

When he was only 19 years old, James Chadwick worked under Ernest Rutherford during his gold foil experiments. More than 20 years later, Chadwick isolated the neutral particle Rutherford had proposed. He called it the neutron.

Experiment Chadwick found the neutron during his experiments in 1932. He and his fellow scientists bombarded beryllium atoms with alpha particles, and an unknown radiation was produced. It was composed of particles with a neutral electrical charge.

Conclusion Chadwick was able to determine that the neutron existed in an atom's nucleus and that its mass was very close to the mass of a proton.

Crookes's Cathode

Thomson's discovery of the electron would not have been possible without the experiments of Sir William Crookes. Crookes created an innovation called a cathode tube, which allowed electrically charged gas particles to flow between electrodes positioned at both ends of a sealed tube.

In one of his experimental investigations, Crookes placed an object within the path of the cathode ray. Turn on the cathode ray below, and observe what happens. What conclusion can be drawn from these results?

Thomson's Electrons

The English physicist Sir J .J. Thomson conducted more experiments using cathode ray tubes. Use the cathode ray below to repeat some of the investigations Thomson conducted.

What do you think will happen if you place a negatively or positively charged rod near the cathode ray tube? What do you think will happen if a different element is used as the negative electrode in the cathode ray tube? Select each rod to view the experiments.

Rutherford's Nucleus

Rutherford expected heavy, fast-moving alpha particles to be able to pass straight through a piece of thin gold foil without being affected by the atoms of gold. He assumed that Thomson's plum pudding model was accurate, that matter and mass were evenly distributed throughout the atom. If this model was correct, there should have been nothing to alter the course of the heavy particles flying through the foil.

Bohr's Quantum Model

Rutherford expected heavy, fast-moving alpha particles to be able to pass straight through a piece of thin gold foil without being affected by the atoms of gold. He assumed that Thomson's plum pudding model was accurate, that matter and mass were evenly distributed throughout the atom. If this model was correct, there should have been nothing to alter the course of the heavy particles flying through the foil.

A Challenge to the Bohr Model

For the most part, Bohr's model has been consistent with current research. However, there is a new model that adds to our understanding of atomic structure and electron behavior. The scientists Schrodinger and Heisenberg developed an alternate theory about atomic nature. Using calculations to determine electron position, they explained that electrons probably do not move in exact orbits like planets in a solar system. Instead, they exist in an electron cloud around the nucleus. They proposed that we can only predict the probable orbital path or position of electrons in an atom. Where the electron cloud is most dense, the probability of finding the electron is greatest.

Parts of the Atom

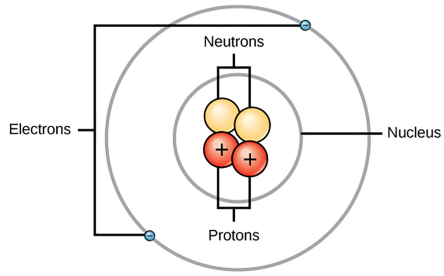

Figure 2.1: Image of the model.

Proton A neutral atom for the element lithium has three protons and three electrons. If one of its electrons happened to be removed from the electron cloud, then it would have three protons and two electrons. This greater number of protons would now make this entire atom a positively charged ion. It would attract negatively charged substances to it.

Electron A neutral atom for the element oxygen has eight protons and eight electrons. If it happened to gain one electron from somewhere, then it would have eight protons and nine electrons. This greater number of electrons would now make this entire atom a negatively charged ion. It would be attracted to positively charged substances.

Neutron Don't be tricked. Neutrons have no charge. If an atom has a greater number of neutrons than electrons and protons, and the numbers of electrons and protons are equal, then the atom is still neutral. It is possible for atoms of the same element to have different numbers of neutrons. These types of atoms are called isotopes. However, this does not affect the charge of the atom, only its mass.

Oct 06 2021 Wen (08:00:23)

Lesson 2: Electromagnetic Radiation **Unit 2**

At one time, matter and energy were considered distinctly different. Matter had mass and volume; energy did not. Instead, energy traveled in a wave. Comparing the two would be like comparing something physical in your hands to something invisible. At least, that's how it seemed.

Electrons behave like waves and particles

Then a scientist named **Max Planck** discovered that energy could have both wavelike and particle-like properties. He also discovered that energy was not continuous, but seemed to transfer in packets. Soon after, Albert Einstein agreed with Planck and added to his hypothesis by proposing that energy transferred in discrete packets called quanta.

Planck's and Einstein's ideas about energy also helped explain the dual nature of electrons. Electrons behave as particles when they orbit in a cloud around the nucleus or when they interact with other substances in a defined position.

To observe electrons behaving as waves, let's look at Feynman's double-slit experiment. In this experiment, a stream of electrons was sent through a wall with two small open slits. The electrons hit the backdrop in specific areas like particles, but when the areas were connected, they showed a wavelike interference pattern. This showed that electrons seemed to have the properties of both particles and waves.

Characteristics of Waves

Radiation is the emission of energy in the form of waves. All waves are repetitive, come in various sizes, and carry energy. When you think of waves, you may think of waves rolling onto a beach. Although they may move in similar ways, an electromagnetic wave is different from ocean waves because they can travel through the emptiness of space.

The lowest point of a wave is a trough, and the highest point of a wave is a crest. The vertical distance between the crest or trough and the center line of the wave is its amplitude. Amplitude measures the height, or intensity, of the wave.

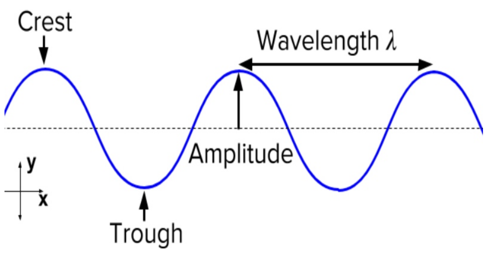

Figure 2.2: Here is a detailed description of the wave.

Electromagnetic Spectrum

- **Radio Waves**: Radio waves are radiation used for radio, TV antennas, and cell phones.
- **Microwaves**: Microwaves are radiation used in radar and to heat food.
- **Infrared Radiation**: Infrared radiation, also known as infrared light, is heat, which is emitted by all objects. The military uses infrared to see how many people are in buildings or to see people at night.
- **Visible Light**: Visible light is all the radiation humans can see.

- **Ultraviolet**: Ultraviolet light is radiation given off by the sun that causes sunburns.
- **X-rays**: X-rays are radiation used in the medical field to look at people's bones and insides. This type of radiation is created when high-velocity electrons collide.
- **Gamma Rays**: Gamma rays are radiation given off by radioactive substances.

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Lesson 3: Quantization of Energy **Unit 2**

> **Definition 65** (Excited State)**.** When elements are heated or energized, their electrons absorb energy and transition to a higher energy level.

Eventually, electrons fall back to a lower orbit or their ground state; but when they do, they release energy in the form of light. This light appears in different colors, depending on the frequency and wavelength of the light released.

When this light is analyzed using a spectroscope, a unique set of colored lines called an emission spectrum identifies the element. This is the emission spectrum for the element hydrogen.

Bohr's model of the hydrogen atom showed the electrons traveling in specific orbits, or paths, around the nucleus.

He suggested that the electrons have a fixed amount of potential energy related to each orbit. The closer the orbit is to the central nucleus, the lower the electron's potential energy.

When you climb a ladder, your potential energy changes depending on your position in relation to the ground.

The higher you climb, the farther you can fall.

This can be compared to the orbits of **Bohr's model**. Each orbit has its own small, definite amount of potential energy; scientists say that this energy is quantized.

Just like the high rungs of a ladder, the higher orbits of an atom are not as stable for an electron.

The electrons eventually fall back to a lower orbit, closer to the nucleus.

Determining the energy of photons

Planck's hypothesis about energy and quanta led to the discovery of the **photon**.

> **Definition 66** (Photon)**.** A photon is a basic particle representing a quantum of energy, E. A photon of light is emitted by an atom when an electron within the atom transitions to a lower energy level.
>
> Conversely, a photon of energy is absorbed by an atom when an electron within the atom moves to a higher energy level. Because energy is transferred in quanta, the photon's entire energy is transferred, not just a fraction of it.

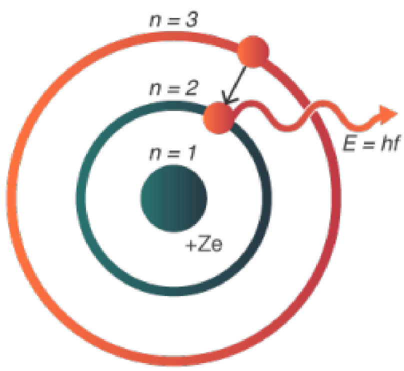

Figure 2.3: Here is what the energy of the photon looks like.

Definition 67 (Planks Constant)**.**

$$E = hf \tag{2.1}$$

The energy of a photon is directly proportional to the frequency of the photon's emission or absorption. Energy is measured in Joules, h is **Planck's constant** of:

$$6.626 \times 10^{-34} \times s$$

and f is the frequency of the energy.

The amount of energy in a photon is influenced by the frequency of its energy source. Examine the electromagnetic spectrum chart below. Notice that as the frequency of energy decreases, the energy of photons from that source also decreases. Because of the inverse relationship between frequency and wavelength, when the wavelength of energy decreases, the energy of photons from that source increases.

Oct 12 2021 Mon (14:38:20)

Lesson 4: Quantum Models **Unit 2**

Once we have determined where an electron could be in the electron cloud, it has already moved on, and the prediction of its location is obsolete. Therefore, we can only describe the probable locations of electrons within an atom's electron cloud.

The probable location of electrons within an atom can be determined using the **Quantum Model**.

Definition 68 (Quantum Model)**.** A mathematical model based on quantum theory that represents the probable state of electrons within an atom.

The probable position of every electron can be described by four quantum numbers. The first number is called the **Principal Quantum Number (n)**. It can be any whole number from one to seven.

The **Principal Quantum Number** refers to the energy levels around the nucleus where electrons are located. These energy levels are also called **Energy Shells**.

Energy levels contain **Sub-shells** that are designated with letters such as

1. s
2. p
3. d
4. f

Due to the number of electrons in most atoms, multiple **Sub-shells** exist within each energy level.

Some **Sub-shells** have multiple orbitals—regions inside an atom where electrons are mostly likely to be found. However, each orbital (no matter the type) can only hold two electrons.

The s orbital The s orbital is spherically shaped. Because the electrons move so quickly, they can be located anywhere inside the sphere.

The s **Sub-shell** contains 1 orbital and is denoted by a line. The electrons are denoted as arrows. An up arrow represents an electron that has a clockwise spin. A down arrow represents an electron with a counterclockwise spin.

The p orbital The p orbital is dumbbell-shaped.

There are 3 of these orbitals, one on each axis.

The p **Sub-shell** has 3 orbitals and is denoted by three lines. The **Sub-shell** can hold a maximum of 6 electrons.

The d orbital The d orbital is nondescript in shape. The d **Sub-shell** has five orbitals and is denoted by 5 lines. This **Sub-shell** can hold a maximum of 10 electrons.

The f orbital The f orbital is nondescript in shape.

The f **Sub-shell** has 7 orbitals and is denoted by seven lines. This **Sub-shell** can hold a maximum of 14 electrons.

A configuration describes the arrangement of objects. Like a map, it shows the location of objects and the distance between them. Similarly, an electron configuration represents the arrangement of electrons around the nucleus of an atom. Each element has its own electron configuration. But unlike a map, an electron configuration doesn't show exact locations for electrons, only their probable placement in energy levels and **Sub-shells**. Remember, within these **Sub-shells** electrons are in constant motion.

All systems in nature tend to take a position or arrangement that has the lowest possible potential energy. The lowest potential energy arrangement for an atom's electrons, called the ground-state electron configuration, is the arrangement that places them closest to the nucleus.

Oct 14 2021 Thu (08:29:41)

Lesson 5: Honors Electrons **Unit 2**

What are the quantum numbers that describe probable electron location?

Definition 69 (Principal Quantum Number (n))**.** The Principal Quantum number describes the energy level within an atom. Positive whole numbers, from one to infinity, are used to represent energy levels. These numbers are used to represent energy levels in electron configurations and orbital notation.

$$1s^2 2s^2 2p^2 \tag{2.2}$$

Definition 70 (Angular Momentum Quantum Number (l))**.** An electron needs both the **principal number** and the **angular momentum quantum number** to identify its specific sublevel. The angular momentum quantum number describes the type of sublevel ($s, p, d, f, \ldots$) using a number instead of a letter ($l = 0, 1, 2, 3, \ldots$).

Basically, electrons in the $2s$ subshell can also be described as

$n = 2, l = 0$ using quantum numbers. It's two different ways of describing the same probable positions for electrons. The principal quantum number $n = 2$ means it's in the second energy level of the atom. The number $l = 0$ tells you that the electron is found in an s sublevel.

Let's use the element chlorine as an example. Chlorine has 17 total electrons. It has an electron configuration of:

$$1s^2 2s^2 2p^6 3s^2 3p^5 \tag{2.3}$$

What are the first 2 quantum numbers for the 6 electrons in the $2p$ subshell?

$$n = 2, l = 1 \tag{2.4}$$

The energy level is 2, and the p subshell is represented by $l = 1$.

Definition 71 (Magnetic Quantum Number (m_l))**.** The **magnetic quantum number (m_l)** represents the specific orbital an electron is likely to occupy. Each type of sublevel has a specific number of orbitals, and each orbital can hold two electrons.

The range of magnetic quantum numbers for a given sublevel (1) is from -1 to $+1$. For example, in an s sublevel ($l = 0$) there is only one available orbital.

It has a **Magnetic Quantum Number** of $m_l = 0$. A p sublevel ($l = 1$) has 3 orbitals with **Magnetic Quantum Numbers** of $m_l = -1, -, +1$.

$$n = 2, l = 1, m_l = -1, 0, +1 \tag{2.5}$$

The magnetic quantum number is $-1, 0$ and $+1$ because the p subshell consists of 3 orbitals that can hold electrons.

Definition 72 (Spin Quantum Number (m_s))**.** The combination of:

- Principal
- Angular Momentum
- Magnetic Quantum Numbers

identifies the location of an electron by:

- Energy Level
- Sublevel
- Orbital

Each orbital can hold 2 electrons, so 2 electrons in the same orbital would share the same values for the first 3 quantum numbers. A $4th$ quantum number, called the **spin quantum number** (m_s), is needed to identify the specific electron within an orbital. It represents the spin direction of an electron on its axis.

According to Hund's second rule, two negatively charged electrons located near each other in the same orbital area are most stable when they spin in opposite directions. There are only 2 choices for the spin quantum number:

$$+\frac{1}{2}\text{OR} - \frac{1}{2} \tag{2.6}$$

An electron can only be assigned 1 spin quantum number. These two values represent the two different directions an electron can spin on an imaginary axis.

Let's give it a shot.

Chlorine has 17 total electrons. It has an electron configuration of:

$$1s^2 2s^2 2p^6 3s^2 3p^5 \tag{2.7}$$

Which quantum numbers represent the probable location and motion of an electron in the $2p$ subshell and orbital $m_i = 0$?

$$n = 2, l = 1, m_l = 0, m_s = +\frac{1}{2}\text{OR} - \frac{1}{2} \tag{2.8}$$

The spin number is $+\frac{1}{2}\text{OR} - \frac{1}{2}$ because $m_l = 0$ orbital holds 2 electrons. The answer could refer to either electron within the orbital. However, the spin quantum number can only be assigned to 1 electron in the orbital, not both. Therefore, the answer could be:

$$n = 2, l = 1, m_l = 0, m_s = +\frac{1}{2} \tag{2.9}$$
$$\text{OR} \tag{2.10}$$
$$n = 2, l = 1, m_l = 0, m_s = -\frac{1}{2} \tag{2.11}$$
$$\tag{2.12}$$

How do orbital notation diagrams model electron arrangements within an atom?

Negatively charged electrons repel each other, so they spread out as much as possible within the orbitals of a given sublevel. This is described by **Hund's first rule**, which stats:

> *The lowest energy atomic state is the one that maximizes the total spin quantum number for the electrons in the open subshell*

This atom has 7 electrons. Each orbital is represented by a line and the energy level and sublevel is shown below the line. Each electron is represented by an arrow. 2 Electrons can be placed in the level $1s$ and $2s$ subshells. Each of the two electrons is represented by an arrow. One is pointed up to show a clockwise spin, and the other is shown pointing down to indicate a counterclockwise spin. According to Hund's second rule, 2 electrons in the same orbital will spin in opposite directions to reduce the repulsion of the electrons.

Figure 2.4: Hund's first rule.

How do the electrostatic forces between electrons influence their orbital spin?

Electric charge is a physical property of matter that causes a particle to experience a force when interacting with other charged particles. These electrostatic forces create an electric field that attracts or repels other charged particles.

The pattern of lines in the images below are called electric field lines. These lines point in the direction another positive charge would follow if placed on these lines.

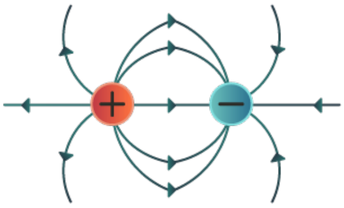

Figure 2.5: Particles with opposite charges attract one another. The electric field lines are directed away from the positively charged particle toward the negatively charged particle.

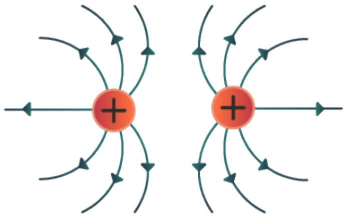

Figure 2.6: Two particles with the same charge repel each other. The electric field lines of each positively charged particle are directed away from one another.

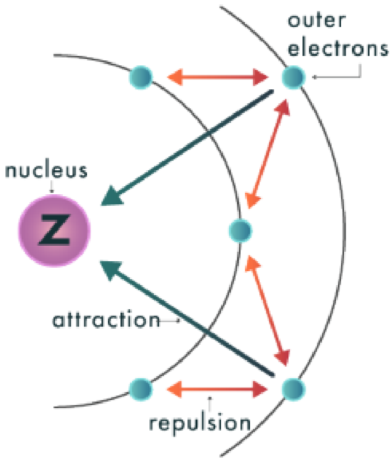

Figure 2.7: This is what it looks like inside the atom.

Electric charge can be neither created nor destroyed. The total amounts of positive charges and negative charges in the universe are always equal, creating a neutral net charge of 0.

Inside the Atom The electrostatic forces between tiny subatomic particles are relatively enormous. The repulsive forces between electrons and the attractive forces between electrons and protons is so significant, it influences the arrangement of electrons within energy levels, subshells, and individual orbitals. It also affects electron spin, preventing two electrons from spinning in the same direction within the same orbital. Instead, electrons within the same orbital spin in opposite directions from one another.

Between Atoms In a neutral atom, the number of positively charged particles (protons) equals the number of negatively charged particles (electrons). Therefore, the net charge on the atom is 0.

If an electron moves from one neutral atom to another, two charged atoms are created. The first atom loses one electron, and so it acquires a $+1$ charge. The second atom gains one electron, so it acquires a -1 charge.

How is the strength of an electrostatic force determined?

The relationship that explains the force between charged particles was discovered by Charles Coulomb in the 18th century. He discovered that electric force decreases inversely as the square of the distance between charges. This is referred to as Coulomb's law.

Basically, as the distance between charged particles decreases, the force between them increases, and vice versa. Also, an increase of electrical charge increases the electrostatic forces between them. A decrease in electrical charge decreases the electrostatic forces between them.

Definition 73 (Coulomb's Law of Electrical Forces)**.** The force between two charges is directly proportional to the product of the charges and inversely proportional to the square of the separation distance. The force between particles acts along a straight line from one charge to the other.

$$F = \frac{kq_1q_2}{d^2} \tag{2.13}$$

F = The electrical force between the two charges, measured in Newtons. k = $8.99 \times 10^9 \frac{Nm^2}{C^2}$. q_1 = Quantity of charge of one particle, measured in coulombs. q_2 = The quantity of charge of the second particle, measured in coulombs. d^2 = The distance between the objects squared, measured in meters.

Nuclear Charge According to Coulomb's law, charged particles that are closer to one another or have a larger product of electrical charge will experience stronger electrostatic forces. This is certainly true when it comes to the interactions between electrons and the nucleus. The net positive charge associated with the protons inside the nucleus of an atom is referred to as its nuclear charge. Electrons closer to the nucleus in energy level n = 1 will have a stronger attraction toward the nucleus due to its nuclear charge. This means it will take more energy to move electrons closest to the nucleus from their orbital regions. Remember, electrons will jump to higher levels with the addition of energy. But those closest to the nucleus have strong electrostatic forces holding their position.

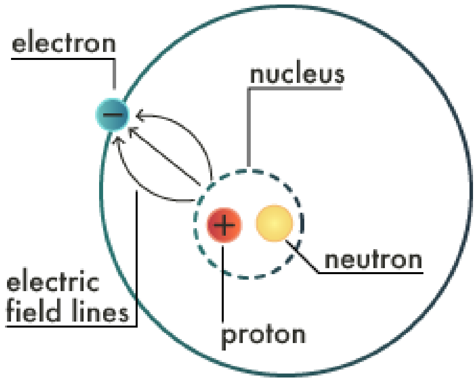

Figure 2.8: This what it looks like when an atom has a nuclear change.

Oct 14 2021 Thu (06:00:23)

Lesson 6: Periodic Table **Unit 2**

There are many ways to sort items. Often, the easiest way to organize things is by their similarities. All items could be grouped with other items of the same

color or **the same type**, such as:

- Plants
- Animals

To group elements, **Mendeleev** compared their:

- Physical properties
- Chemical properties

and searched for common traits.

History of the Periodic Tabel

It's been more than 200 years since the Periodic Table of elements was first described by Dmitri Mendeleev, but the idea of an element as a basic type of material must have occurred to early humans as they selected certain types of rocks for tools and later melted metal into weapons and baked wheat into bread.

Greek scholar **Democritus** thought that substances were made up of tiny invisible particles that could not be broken into smaller pieces.

Well, atoms—"atomos" in Greek means indivisible. And the way he defined it was by imagining, say, a gold wire that you rolled incredibly thin, and you chopped it in half and chopped it in half again, and then again, and then again. Eventually you would get to something so small that you couldn't chop it in half any more, and it would still be gold. That is an indivisible unit, an atom of gold.

Other ancient Greek scholars classified the fundamental elements as:

- Fire
- Earth
- Air
- Water

The philosopher Aristotle believed that these elements could be mixed to create new materials. This sparked a centuries-long interest in alchemy, specifically transmuting ordinary metals like lead into gold.

Scottish alchemist Robert Boyle concluded in the late 1600s, however, that although metals like gold and lead are both made of atoms, the size and shape of the atoms are different.

Nearly 100 years later, French chemist Antoine Laurent de Lavoisier offered the first modern definition of an element, a chemical substance that could not be broken down into another substance. He named more than 30 elements himself.

His contemporary, John Dalton, figured out a way to estimate the weight of each kind of atom. Hydrogen always made up 15 percent of the weight of water, whereas oxygen accounted for 85 percent of water's weight. Using these numbers, Dalton was able to calculate the atomic weight of these two elements.

The list of elements and their atomic weights grew throughout the $18th$ and early $19th$ centuries. But it remained just a list arranged from lightest to heaviest. However, some scientists started to recognize patterns within the list.

There are a couple of patterns detected in the lists of elements. The first are triads, where you take three elements in a row. The atomic weight of the middle element would be the average of the atomic weight of the lighter and the heavier.

Similarly, other chemists noticed that there were natural families in which certain elements grouped. The halogens, for example, which are salt-making elements, usually very highly reactive, tended to have very similar properties, form very similar crystals, and so on. And those were grouped together as a natural family.

Dmitri Mendeleev was writing a new textbook for his students when he noticed a repeating, or periodic, pattern among the 63 known elements.

His periodic table made its debut in 1869. His table looks a little different from the ones that are hanging in chemistry classrooms. The atomic weight increases from top to bottom, and the families run from left to right. So you would have to rotate it clockwise by 90 degrees and then reflect it in the mirror to make it look like our current one.

But it still has the basic properties of gradual increase of atomic weight and periodic repetition of properties across families as you increase the weight. Mendeleev put question marks in spaces where he thought there should be an element of a certain atomic weight. Even though no element of that weight had been discovered yet, he predicted their atomic weights and other features such as their melting and boiling points.

In this sense, Mendeleev did not invent the periodic table of elements. He discovered it. His carefully stacked rows and columns turned a simple list into a snapshot of how matter is organized on Earth and throughout the universe.

Mendeleev looked for similarities between elements to group them together. He started by grouping elements that were metals and nonmetals based on their physical properties.

Physical Properties of Metals:

- Good conductors of electricity and heat

- Solid at room temperature
- Malleable, flexible, and ductile
- Lustrous (shiny)
- Higher density

Physical Properties of Nonmetals:

- Poor conductors of electricity and heat
- Solid, liquid, or gas at room temperature
- Solids are brittle and break easily
- Dull (not shiny)
- Lower density

If an element did not fully fit either grouping, it went to a group called metalloids, which had characteristics of both metals and nonmetals.

Reading the periodic table takes a little practice. The letter or letters that symbolize the element are sometimes self-evident, like H for hydrogen. Some are not so obvious, such as Sb for antimony.

The atomic number at the top of the box represents the number of protons contained in an atom of that element. Underneath the element's name is the atomic weight. Each row of the table is called a period.

All of the elements in a period have the same number of electron shells. Columns in the table are called groups. For most groups the elements have the same number of electrons in their outermost electron shell.

Another way to look at the table is to divide the elements into metals, nonmetals, and metalloids. Most of the elements in the table are metals. When people are asked to name any element, the chances are good that they will come up with the name of a transition metal.

Transition metals are found in numerous products such as:

- Coins
- Jewelry
- Light
- Bulbs
- Cars

Transition metals are called the bridge of the periodic table and include groups 3 through 12.

Group 1 elements, beginning with *lithium* and running vertically to *francium*, are called *alkali metals.*

Group 2 elements, beginning with beryllium and running vertically to radium, are called alkaline earth metals. Elements in these first two groups of the periodic table are some of the most reactive elements known.

In fact, many of the elements are so reactive that water, or in some cases air, can cause them to explode or catch fire. Although many of their names are familiar:

- Sodium
- Potassium
- Calcium

for instance—it is rare to find these elements on their own. Instead, these elements are usually combined with another, more stable element to create a compound.

The elements in *groups* 13 through 16 of the periodic table are a mixed bag. Some of the elements are metals, though they barely act like metals in chemical reactions.

Other elements within these groups are definitely nonmetals, whereas still other elements—metalloids—have a combination of metal and nonmetal features. Scientists find it difficult to agree on one name for all these groups, so many researchers call them the BCNOs after the first element in each group—boron, carbon, nitrogen, and oxygen. BCNOs, depending on whether they act more like a metal or a nonmetal, may give up electrons or share electrons with other atoms. The valence shells of the BCNOs are about half full.

Definition 74 (Atomic Mass)**.** The atomic mass, which is measured in atomic mass units (amu), is an average of the masses of all the isotopes of that element.

Most of the mass of an atom comes from the protons and neutrons. The electrons are so small they don't add much mass to the total.

Definition 75 (Atomic Number)**.** Elements of the periodic table are arranged in order by their atomic numbers. The atomic number, which is usually found above the element's chemical symbol, indicates the number of protons in a neutral atom. Since positive charges equal negative charges in a neutral atom, the number of protons equals the number of electrons.

Definition 76 (Atomic Symbol)**.** Chemical symbols are shorthand versions of the full names of elements. For each element, the chemical symbol is usually derived from the first letter or two of the element name. However, in some instances, it does not follow that pattern.

The number of:

- Protons
- Neutrons
- Electrons

within an atom affects its properties. You can determine the number of each of these subatomic particles when given the atomic number and mass number for a given atom.

Calculating Average Atomic Mass

Differeing Neutrons Isotopes of an element have the same number of protons but a different number of neutrons. This means isotopes of the same element will have different atomic masses.

Carbon has three isotopes: carbon-12, carbon-13, and carbon-14. All three isotopes have six protons, but their neutron numbers are six, seven, and eight. This means that the three isotopes have different atomic masses but share the same atomic number of six.

The number added to the element's name refers to the mass number. So, carbon-13 is an isotope of carbon with six protons and seven neutrons, and carbon-14 is an isotope of carbon with six protons and eight neutrons.

Differeing Masses Most elements found in nature occur as a mixture of two or more isotopes. However, the percentage of each isotope in nature is different. Some isotopes are more common than others.

To calculate the average atomic mass of an element, the percent abundance of each isotope of that element is multiplied by its mass. Then all the atomic masses are added together to create a weighted average that accounts for the percentage of each isotope in nature.

(percent abundance × mass) + (continue for each isotope) = average atomic mass

(Percent Abundance * mass) + Continue for each Isotop = Average Atomic Mass

Isotope Reps There are two main ways to represent an isotope. Simply place a dash to the right of the name with the mass number, such as uranium-235.

Or, place the mass number as a superscript to the upper left of the chemical symbol, and place its atomic number as a subscript to the lower left of the chemical symbol.

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Lesson 7: Periodic Trends **Unit 2**

Periodic Trends influenced by the Nuclear Charge.

As the atomic number increases, the number of protons increases, too. This trend is useful when searching for an element on the periodic table. More massive elements (those with more protons and neutrons) will be toward the bottom of the periodic table, and less massive elements will be toward the top. But this is not the only useful pattern on the periodic table. The periodic table shows many trends related to the properties of elements and how they interact with one another.

Effective Nuclear Charge An atom with positively-charged nucleus surrounded by an electron cloud. The electrons closest to the nucleus provide a shield for the outer shell electron.

Charged particles like electrons are affected by electrostatic forces. Negatively charged electrons repel one another, and protons $(+)$ and electrons $(-)$ attract each other. So, the relative positions of electrons within an atom are influenced by the positive charge of the nucleus.

But not all electrons **"feel"** the full positive charge of the nucleus. Electrons farthest from the nucleus are partially shielded by the electrons in the energy shells closer to the nucleus. So, outer shell electrons feel an effective nuclear charge (z_{eff}) that considers the shielding effect of the inner core electrons.

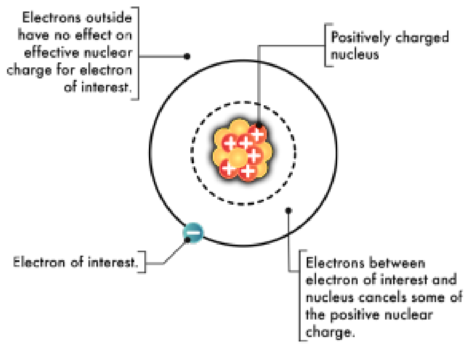

Figure 2.9: In-detail look at an atom with a nuclear charge.

Screening Constant In general, subtracting the number of shielding (core) electrons (also called the screening constant) from the total nuclear charge (number of protons) gives an estimate of the effective nuclear charge felt by the electrons in an atom's outermost shell.

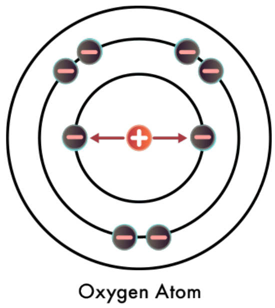

Figure 2.10: Removing electrons.

Estimating Effective Nuclear Charge The $+8$ charge of oxygen's nucleus is shielded by the two core electrons in the first energy level. This means each of the outer shell electrons in oxygen feels an estimated effective nuclear charge of about $+6$. This is calculated by:

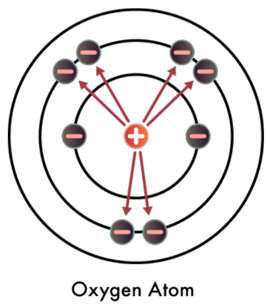

Figure 2.11: Here is how we estimate the effective nuclear charge.

$$8 - 2 = +6 \tag{2.14}$$

$$\text{(total nuclear charge) - (shielding electrons) = (effective charge)} \tag{2.15}$$

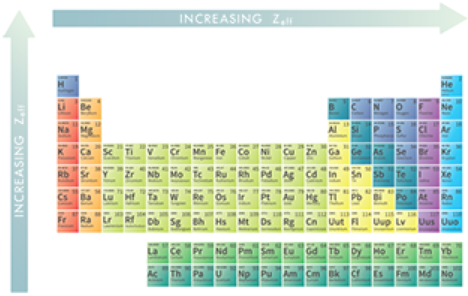

Figure 2.12: Periodic Table.

Effective Nuclear Charge Across Periods Across a period:

- The overall nuclear charge increases by one with each additional proton.
- Each atom also has an additional electron in its outer energy level, but the number of core electrons in the lower energy levels remains the same.
- The screening constant stays the same.

For instance, phosphorous is to the left of sulfur. It has 15 protons and 15 electrons. Like sulfur, it has the same number of electrons in its innermost energy shells, which is 10.

$$\text{Sulfur: } z_{eff} = +16 - 10 = +6 \tag{2.16}$$
$$\text{Phosphorus: } z_{eff} = +15 - 10 = +5 \tag{2.17}$$

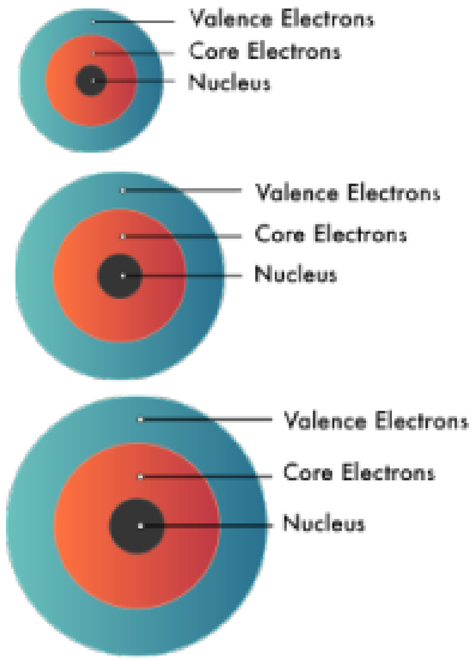

Figure 2.13: Electrons in each group.

Effective Nuclear Charge Down Groups The effective nuclear charge felt by an atom's outer shell electrons moving down a group is affected by competing components.

- Nucleus
- Shielding core Electrons
- Outer Shell Electrons
- Protons increase down a group, causing a greater attraction between protons and outer shell electrons. Electrons are pulled inward toward the nucleus.
- The size of atoms and the number of core electrons also increase down a group, thus increasing the number of shielding electrons.

Therefore, the nuclear charge of outer shell electrons stays relatively constant for each element down a group. This is because the number of additional protons in each element's nucleus is equal to the number of additional shielding electrons in the inner energy levels of the atom.

For example, **Lithium (Li)**, **Sodium (Na)** and **Potassium (K)** in *Group 1* each have an effective nuclear charge of $+1$ felt by its one outer shell electron.

CHAPTER THREE

Molecules and Compounds

Unit 3

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Lesson 1: Valence Electrons **Unit 3**

Factors that Influence Chemical Bonding

Compounds are held together by **Electrostatic Forces** between atoms in a **Chemical Bond**. All chemical bonds involve **Valence Electrons**, but bonds are classified by the way electrons are distributed within them. Here are the two types of bonds:

Ionic Bonds An **Ionic Bond** is a chemical bond that results from the **Electrostatic Attraction** between positive $(+)$ and negative $(-)$ ions. This means that in an ionic bond, electrons are given up by one atom and gained by another atom, and then those atoms are attracted to each other.

Covalent Bonds In a **Covalent Bond**, electrons are shared between two atoms, neither atom completely gaining or losing electrons. Because atoms in **Covalent Bonds** overlap their outer energy levels, both atoms *"Own"* the shared electrons.

In general, atoms of **nonmetals** form **covalent bonds** with each other, and **metal atoms** form **ionic bonds** with **nonmetal atoms**. There are many exceptions to this generalization, but it still helps us predict which elements will bond together and what type of bond they will form.

Valence Electrons Identified in Electron Configurations

The most important electrons in an atom are the ones in its **Outermost** or **valence shell**. The **outermost shell** is the highest energy level that contains electrons in an atom. The valence electrons within the **outermost shell** can be removed from one atom and given to another atom. These electrons can also be shared with atoms of other elements. This is how atoms undergo

chemical reactions or bond with other atoms. For the element krypton, the highest energy level used is level 4. There are two occupied **subshells** in level 4—**subshells** $4s$ and $4p$. The s **subshell** has 2 electrons, and the p **subshell** has 6 electrons. This totals to 8 **valence electrons** for krypton.

Using the Periodic Table to determine an Element's Valence Electrons

The **Octet Rule** states that atoms for elements in the main element groups prefer to have up to 8 **valence electrons**. This fits how the periodic table is organized, with 8 tall columns and several shorter columns for **transition metals**. If the **valence electrons** equal 8, the element is inert, which means it doesn't react with other substances. If an atom of an element has fewer than 8 electrons, it may seek to form chemical reactions with other atoms to get a full outer shell of valence electrons.

Columns 1 **and** 2 All the elements in column 1 have 1 valence electron in their **outermost shell**. For column 2, it's the same idea. Atoms of elements in this column have 2 **valence electrons**.

Transition Metals The transition metals follow unique rules, so their columns cannot directly tell us about their total **valence electrons**. Scientists determine their valence electrons in other ways.

Columns 3 **to** 7 Just like columns one and two, taller columns three through seven indicate the number of valence electrons for the elements within them. For example, elements in column three have three electrons.

Column 8 With the exception of helium, which only needs 2 electrons to fill its **outermost shell**, all other elements in column 8 have 8 **valence electrons**.

Element's Valence Electrons relate to its Chemical Reactivity

Valence electrons are often found in the s and p sublevels of the outermost energy level in an atom. Each subshell has its own unique shape. Subshells d and f have some challenges to bonding that the s and p subshells do not have. That is why d and f electrons are not usually included in the valence and are rarely involved in chemical reactions.

Empty spaces in the subshells of the outermost shell also affect reactivity. Atoms prefer eight valence electrons. Therefore, they will bond with other substances to gain a full eight. Some atoms may gain electrons to make eight, while others can give electrons away. Let's see how that works.

Valence Electrons affect the Bonding Properties of Elements

Elements with the same number of valence electrons show similar chemical behavior, which is why they are grouped together on the periodic table. These elements also follow similar periodic trends when it comes to ionization energy and electronegativity.

Chemical bonds can range in strength and properties between completely ionic and completely covalent, depending on how strongly the atoms attract the electrons. Comparing their electronegativity values helps chemists estimate the degree of ionic or covalent properties of various chemical bonds. The greater the difference in their electronegativity values, the more ionic the chemical bond between two atoms.

Shorthand Notation to Illustrate Valence Electrons

Full **Electron Configurations** can get very long when atoms have large atomic numbers. Sometimes, the chemist is only concerned with an atom's valence electrons, since those are the electrons that are more involved with bonding and other chemical properties. For this cases, scientists have created a special notation, called **Gas Notation**, for abbreviating the electron configuration.

For example, for **Chlorine (Cl)**, the electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^5$. Using the **Gas Notation**, we can write the electron configuration as **[Ne]** $3s^2 3p^5$

How to Determine the Electron Configuration for Mercury (Hg)

1. Write the symbol of the last element from the previous row in brackets. [Xe] 2. Mercury's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 5p^6 6s^2 4f^{14} 5d^{14}$. Write the rest of the element's notation to complete that element's shorthand notation:

$$\text{[Xe]}6s^2 4f^{14} 5d^{10} \tag{3.1}$$

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Lesson 2: Ionic Bonding **Unit 3**

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Lesson 3: Covalent Bonding **Unit 3**

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Lesson 4: Nomenclature **Unit 3**

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Lesson 5: Molecular Structure **Unit 3**

Electrostatic Forces

Valence shell electron pair repulsion (VSEPR) theory is used to predict the geometry and shape of a molecule, which plays an important role in molecular interactions. This theory is based on the concept that valence electron pairs, whether bonded or nonbonded (lone pairs), repel one another. Lone pairs and bonded pairs will spread out and position themselves around the central atom so they are as far away from each other as possible.

The electron geometry (also called VSEPR shape or VSEPR geometry) of a molecule is determined by the number of **Bonded Electrons** and **Lone Pair Electrons** around the **Central Atom**. **Molecular Shape**, also called the **Molecule Structure**, is a general shape of the molecule that considers the repulsive and attractive forces between electron pairs. More often than not, the **VSEPR** shape and molecule structure are different for a molecule. Two things to remember:

- **VSEPR** geometry includes unshared electron pairs, if any, on the central atom that determines the geometry.
- The molecule structure is the atoms only and the shape the atoms make when bonded together.

Electron pairs exist in many positions, or domains, around a central atom. The number of domains is the sum of bonded atoms and lone pairs of electrons surrounding the central atom.

Polarity of Molecules

Molecular Polarity is dependent on the difference in **Electronegativity** between atoms in a compound and the molecular asymmetry of the compound's structure. Polarity also influences other physical properties, such as melting and boiling points, solubility, and the interactions between molecules.

Nonpolar molecules have an even distribution of charge due to all **Nonpolar Bonds**, or **Symmetrical Polar Bonds** that "cancel out". **Polar Molecules** have a partial negative end and a partial positive end due to asymmetric arrangement of polar bonds.

Lesson 4: Forces and Bonds **Unit 2**

Lesson 7: Honors Organic Chemistry **Unit 3**

CHAPTER FOUR

Reactions

Unit 4

Lesson 1: Conservation of Mass **Unit 4**

Lesson 2: Synthesis and Decomposition Reactions **Unit 4**

Lesson 3: Single and Double Replacement Reactions **Unit 4**

Lesson 4: Combustion and Redox Reactions **Unit 4**

Lesson 5: Honors Oxidation Reduction **Unit 4**

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Lesson 6: Reactions in Our World **Unit 4**

In **Chemical Reactions**, the loss or gain of **Electrons** does not change an element's identity. But during a **Nuclear Reaction**, a **Nucleus** is split apart or 2 smaller **Nuclei** are combined.

In chemical reactions, the loss or gain of electrons does not change an element's identity. But during a nuclear reaction, a nucleus is split apart or two smaller nuclei are combined. This changes the identity of the element because protons and neutrons are gained or lost by the atom over the course of the nuclear reaction. These "lost" protons and neutrons don't just disappear, they become new particles.

Lesson 7: Honors Radioactive Decay **Unit 4**