

Nuclear Model of the Atom

February-11-15 5:56 PM

Formulas & Theorems Covered Today:



Homework:



Notes:

The Nuclear Model of the Atom

- Towards the close of the 19th century, chemists had two tools to help them understand matter
 1. John Daltons Atomic Theory - he described the behaviour of matter in chemical reactions
 2. Dimitri Medeleev's Periodic Table - which listed the known elements (not all had been discovered) in order of increasing atomic mass and repeating properties (later grouped into families)
- We known now that these are not the most current models of each
- The atom for example is not an indestructible particle, it is made up of sub atomic particles, and those are even made of smaller particles
- Chemists however needed Dalton's atomic theory to advance their understanding of matter and its behaviour during chemical reactions
- Dalton is known as the father of modern chemistry
- His theory pushed the world from an understanding that stated that everything was made of 4 elements (earth, fire, water and air) to a world of limitless particles, each with their own properties and each with their own masses
- The idea of this was not all his own, in fact, many scientists before him played an integral role in breaking the dogma that was the make up of the universe
- The issue with Dalton's model however, was that it could predict the formation of simple compounds like CO_2 , and SO_3 , but it ran into issues with more complex compounds because it

had difficulty predicting the ratios in which atoms bond

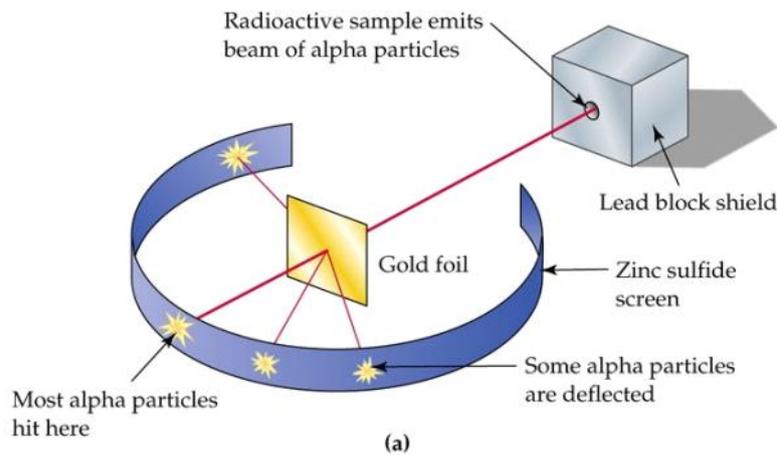
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Discovering The Electron

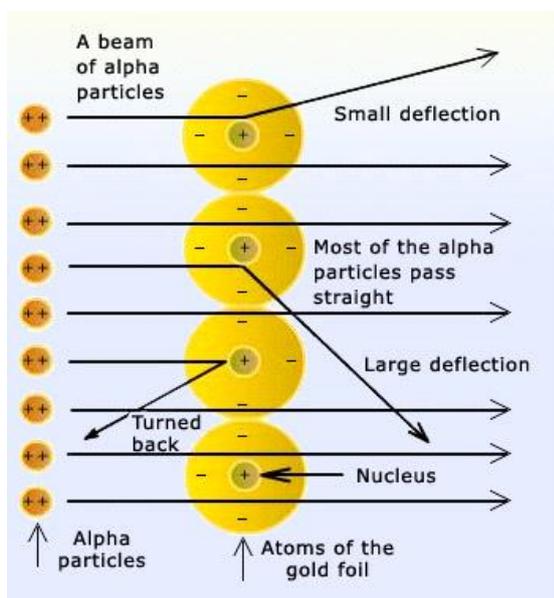
- In 1897 Dalton's atomic theory was shattered by a scientist named Joseph John Thomson, or J.J. Thomson for short
- His discovery of the existence of a negatively of a negatively charged particle with a mass $\frac{1}{1000}$ that of a hydrogen atom
- This particle was called the electron (actual mass $\frac{1}{1837}$ of a hydrogen atom)
- After his discovery of the electron, he developed a new model of the atom which took Dalton's cue ball model of the atom and added into it the electrons and a positive core
- He coined the term "Plum Pudding Model" of the atom which had a positive core and negative particles embedded into it

Gold Foil Experiment

- It was not uncommon for the time that when a new atomic theory came out, other scientists would test, and re-test to see whether or not a theory would produce the same results
- A few years after J.J. Thomson discovered the electron, his student, Ernest Rutherford, created an experiment to test Thomson's theory
- He postulated that if he took a radioactive source, placed in front of a large, thin sheet of gold foil that the majority of the α -particles (fired from the radioactive source) would travel through the gold foil unimpeded
- The thought behind this was that the gold atoms are so large, and thus the spaces between the atoms are so large, that the tiny α -particles (helium nuclei, 2P, 2N) would travel right through
- The radioactive source was placed in a lead block shield and directed towards the thin gold foil
- The room had a zinc sulfide screen (which would flash when exposed to the α -particles) that was placed on a track so that the screen could be positioned at various spots around the room
- He asked his grad students to sit and watch for the small flashes of light caused by the interaction of the α -particles with the zinc sulfide screen
- After seeing that a huge number of the α -particles were going through, Rutherford asked his students to start moving the screen around the room to see if any of the α -particles were being deflected to other locations
- This stroke of genius gave Rutherford and his team a set of data that would change the modern atomic theory once again



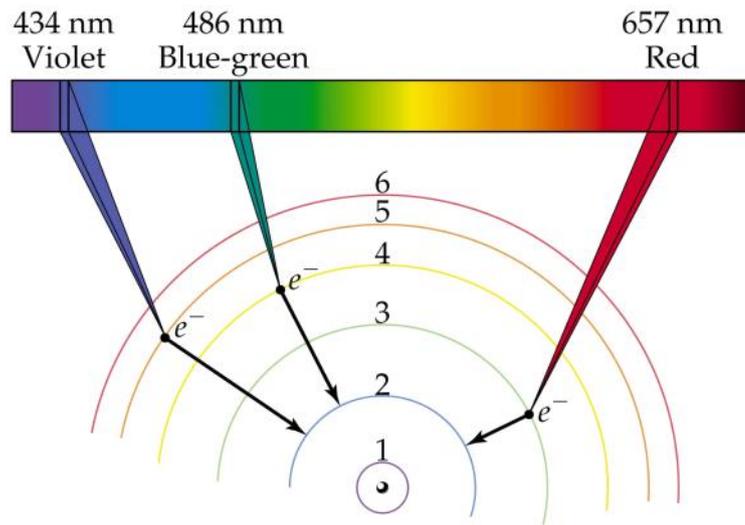
This allowed Rutherford to make the following discoveries:



- The Rutherford model of the atom was coined the planetary model of the atom

Energy Levels and the Bohr-Rutherford Model

- Neils Bohr, one of Rutherford's students, furthered the atomic model with his study on emission spectra
- He found that when a hydrogen atom had its electrons excited, they would jump to unoccupied energy levels
- This increase in potential energy is short lived and the electrons want to ultimately be in the lowest energy level
- In order to achieve that, the excited electrons (excited state) would drop back down to the lower energy levels
- The potential energy that they had gained from being excited would then be emitted in the form of light as per the following diagram



- This, accompanied with Rutherford's discovery of the nucleus became the dominant atomic model until we move into quantum mechanics

Summary

After the lecture, summarize the main points of this lecture topic.

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The Quantum Mechanical Model of the Atom

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Formulas & Theorems Covered Today:



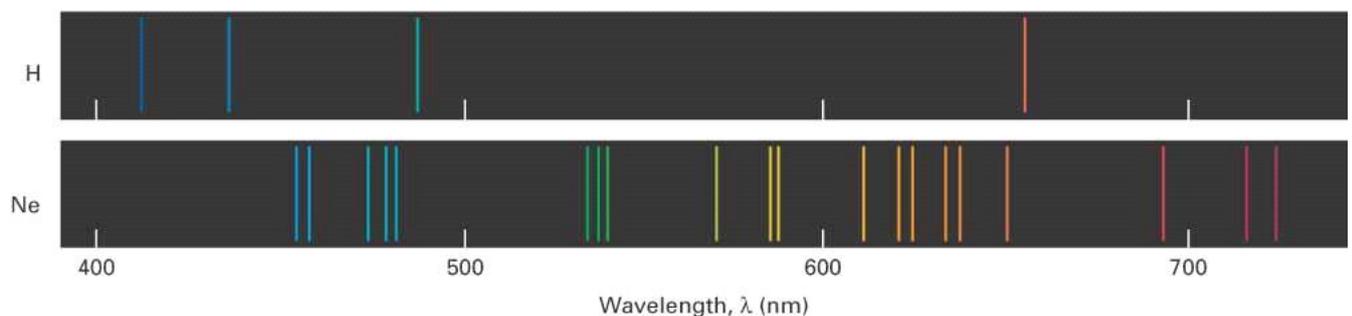
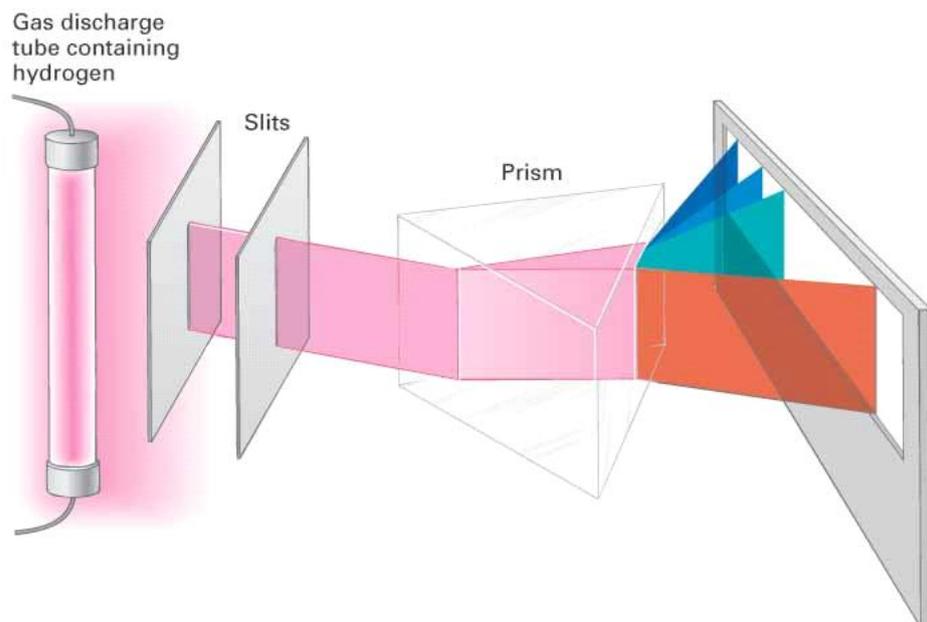
Homework:



Notes:

The Quantum Mechanical Model of the Atom

- Bohr's Model explains the emission spectrum for hydrogen but other atoms posed a problem when it came to explaining their emission spectra
- If you look at the emission spectra of other elements, like Ne (seen below), they have more spectral lines than hydrogen



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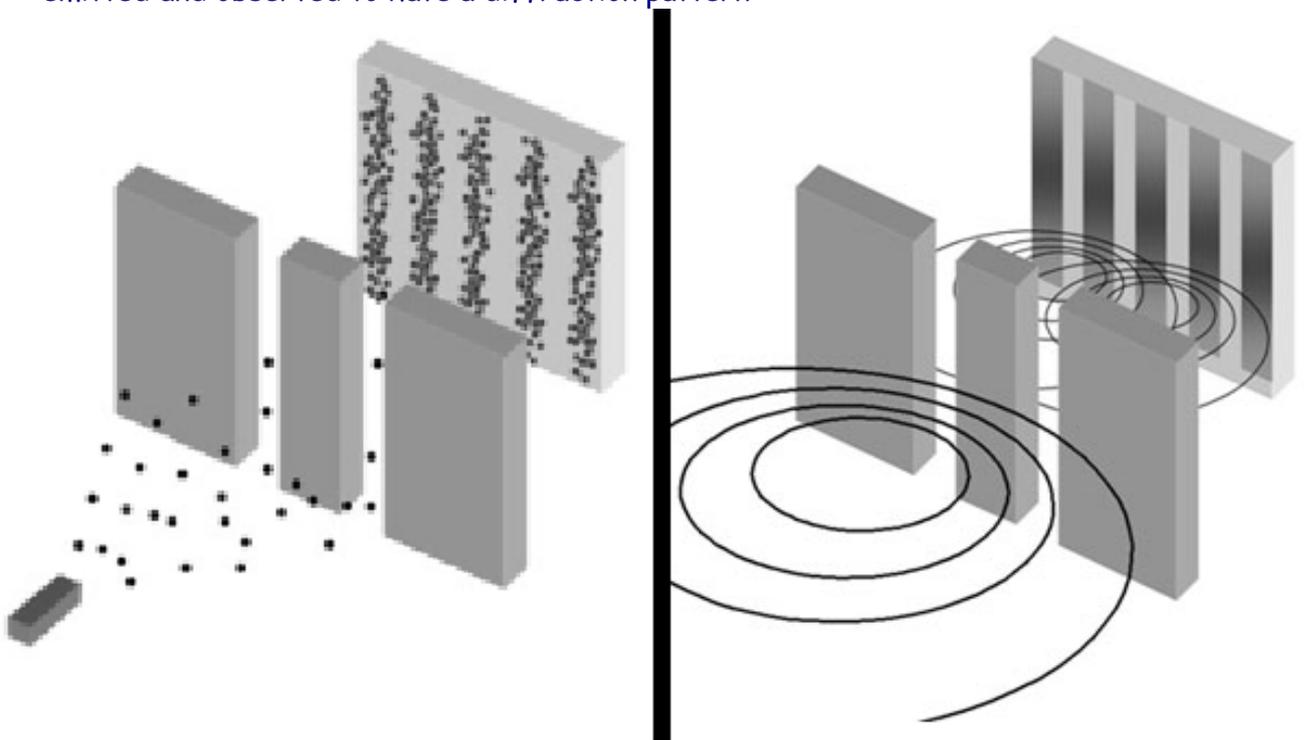
- The reason for the significant increase in the number of lines that are present is the number of electrons that are found in the elements
- Bohr's simplistic example of energy levels was not enough to explain these added lines and so

scientists had to reconsider the idea of energy levels

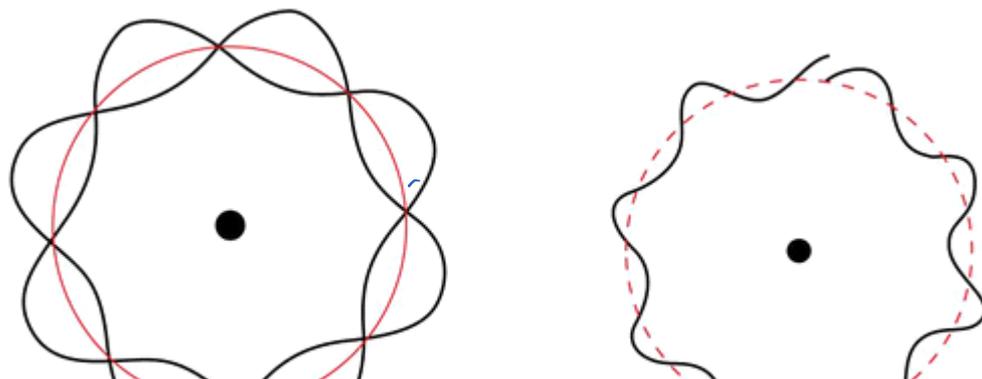
- In the hydrogen spectrum, it is easy to distinguish between the bands so the energy levels could be easily distinguished (as seen above)
- Large elements, like Neon, certainly have the major energy levels present, but there are other bands also present that represent sub energy levels within those major energy levels

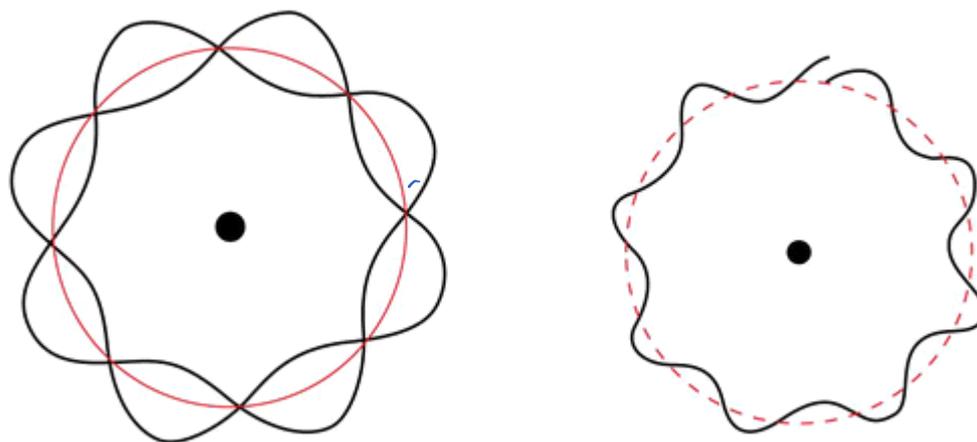
The Discovery of Matter Waves

- It was general knowledge that energy had matter like properties but in 1924, a physics student name Louis DeBroglie speculated that matter had wave like properties
- He created an equation that allowed him to calculate the wavelength that was associated with any object
- Large objects, like a baseball, has a very small wavelength, λ , associated with it relative to their size, but electrons, relative to their size have a huge λ
- This hypothesis has been profound effects on motion and matter all around us
- His hypothesis was later tested and confirmed in 1927 when streams of electrons were being emitted and observed to have a diffraction pattern



- The discovery that electrons possessed certain amounts of energy and acted like waves is quite important and profound to all of science
- It explains why electrons don't simply crash into the nucleus of an atom
- They cannot crash into the nucleus because it is not allowed based on the smallest energy wavelength and the idea of standing waves as shown in the diagram below





Quantum Model

- Erwin Schrödinger took DeBroglie's idea of matter waves, coupled with Einstein's idea of quantized energy (photons) and amalgamated the two into a single unified equation
- This equation, coupled with Heisenberg's uncertainty principle was the birth of quantum mechanics
- Quantum Mechanics is used to discuss the wave like properties of matter, and is quite useful because it has testable predictions and the predictions agree with the hypotheses

Heisenberg's Uncertainty Principle

- He proposed that it is impossible to know the position and momentum of any object beyond a certain measure of precision
- Electrons are so small, that the measure of precision comes into play and so the electrons position around a nucleus, and its momentum, cannot be simultaneously known
- If you know a lot about its position, then you know nothing about its momentum, and if you know a lot about its momentum, then you know nothing about its position
- The simple act of looking changes these two quantities
- We have statistics that will help us predict the likelihood of finding an electron at a particular point around the nucleus of an atom
- To do this, we can use the wave equation to tell us information about the electrons energy and location within an atom
- As chemists, we call these atomic orbitals

Summary

After the lecture, summarize the main points of this lecture topic.

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Quantum Numbers and Atomic Orbitals

Formulas & Theorems Covered Today:



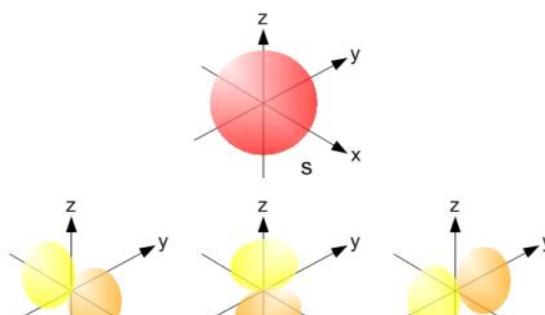
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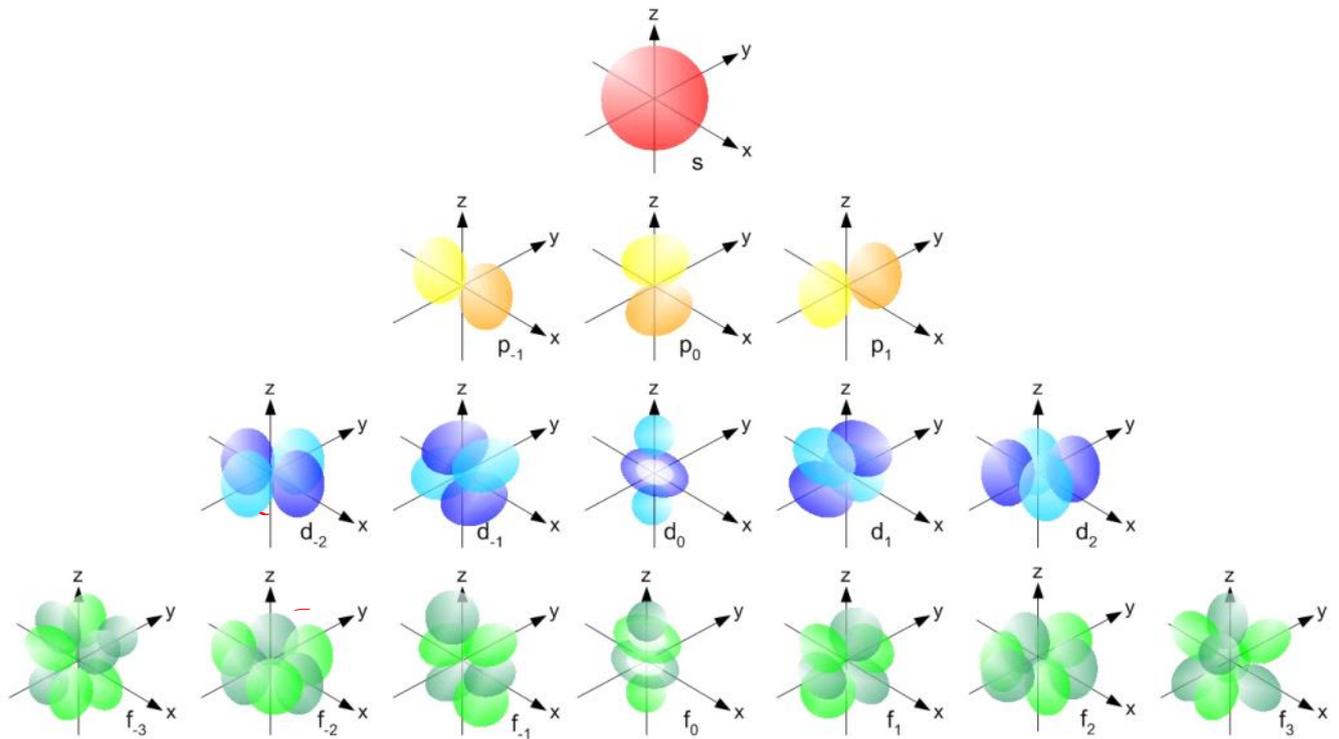


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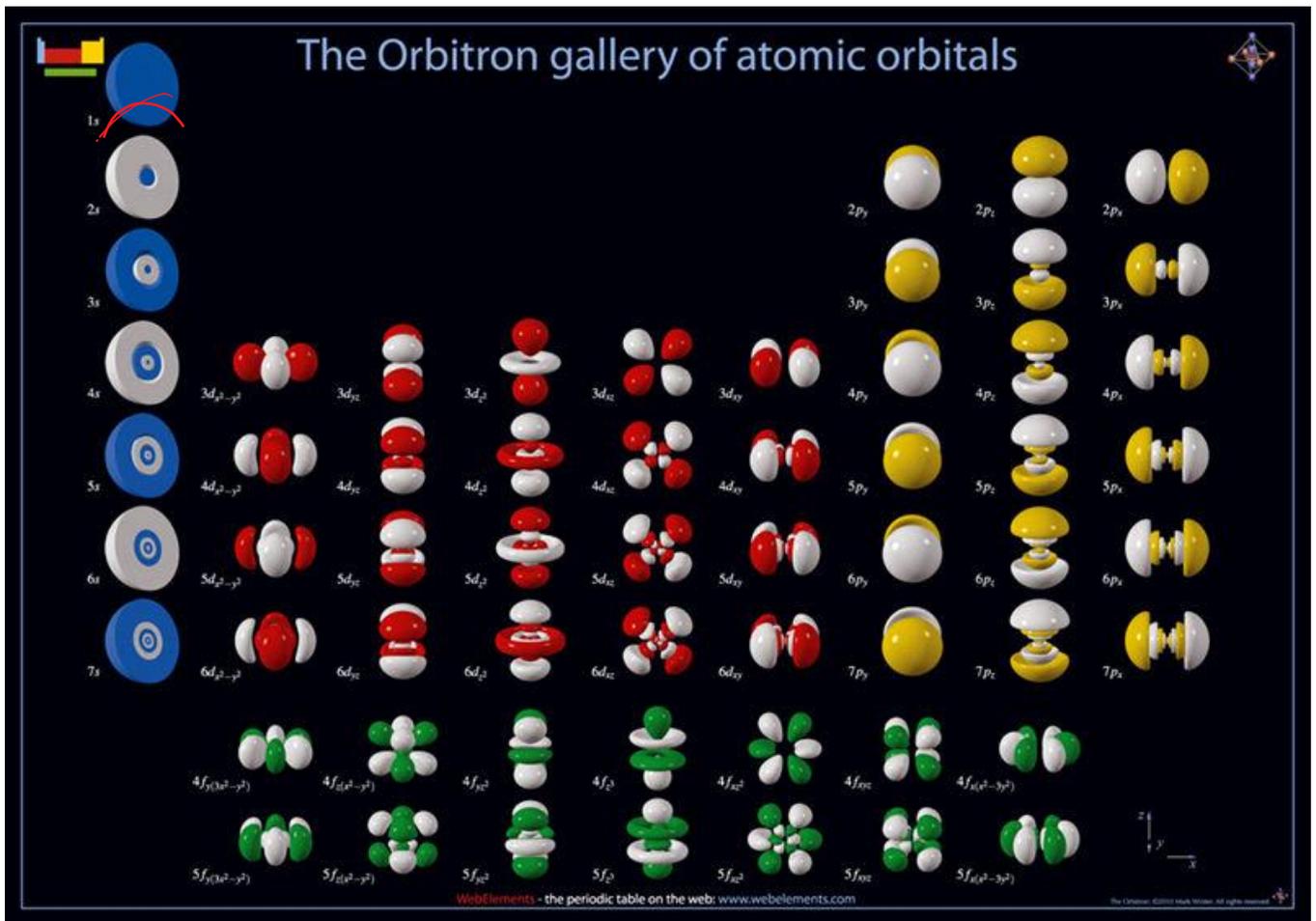
Quantum Numbers and Orbitals

- The most stable state for an atom is called its ground state
- There are 4 quantum numbers that are assigned to any electron within an atom to give us information about that element (namely the energy it possesses and the location of that element)
 - **The Principle Quantum Number - The first Quantum Number**
 - The principle quantum number is denoted with the letter "n" and can be any positive whole number
 - It specifies the size and energy of an orbital
 - The larger the number, the higher the probability of finding that electron further from the nucleus
 - The greatest number of electrons that can be found in any given atomic energy level is given by $2n^2$
 - **The Azimuthal Quantum Number - The second Quantum Number**
 - This number is denoted with a script "l" and can range from 0 --> n-1
 - This quantum number represents the shape of the atomic orbital
 - $l = 0$ --> denotes the "s" orbital
 - $l = 1$ --> denotes the "p" orbital
 - $l = 2$ --> denotes the "d" orbital
 - $l = 3$ --> denotes the "f" orbital





- **The Magnetic Quantum Number - The third Quantum Number**
 - This number is denoted by "m_l" and has a value that can range from -l → l
 - This number specifies which sub orbital within the major atomic orbitals that the electron is found
- **The Magnetic Spin Quantum Number - The fourth Quantum Number**
 - This is called the spin quantum number and is denoted by "m_s"
 - This can have a value of $+\frac{1}{2}$ or $-\frac{1}{2}$
 - The Pauli Exclusion principle states that no two electrons can possess the same 4 quantum numbers as they are not allowed to occupy the same energy at the same time



Electron Configurations

Formulas & Theorems Covered Today:



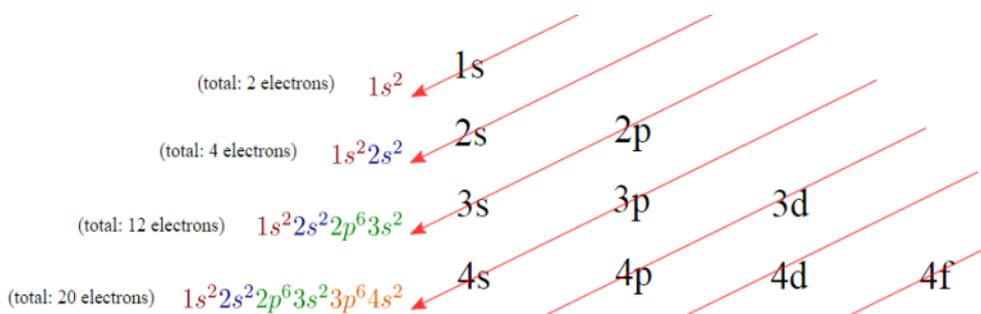
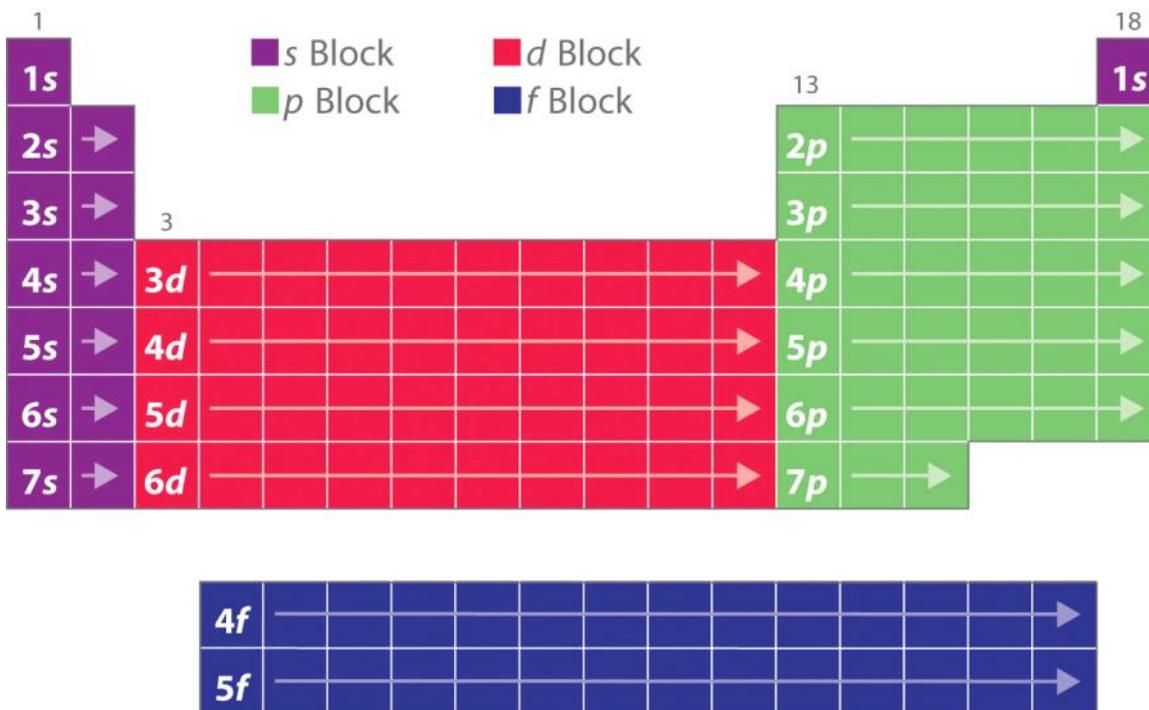
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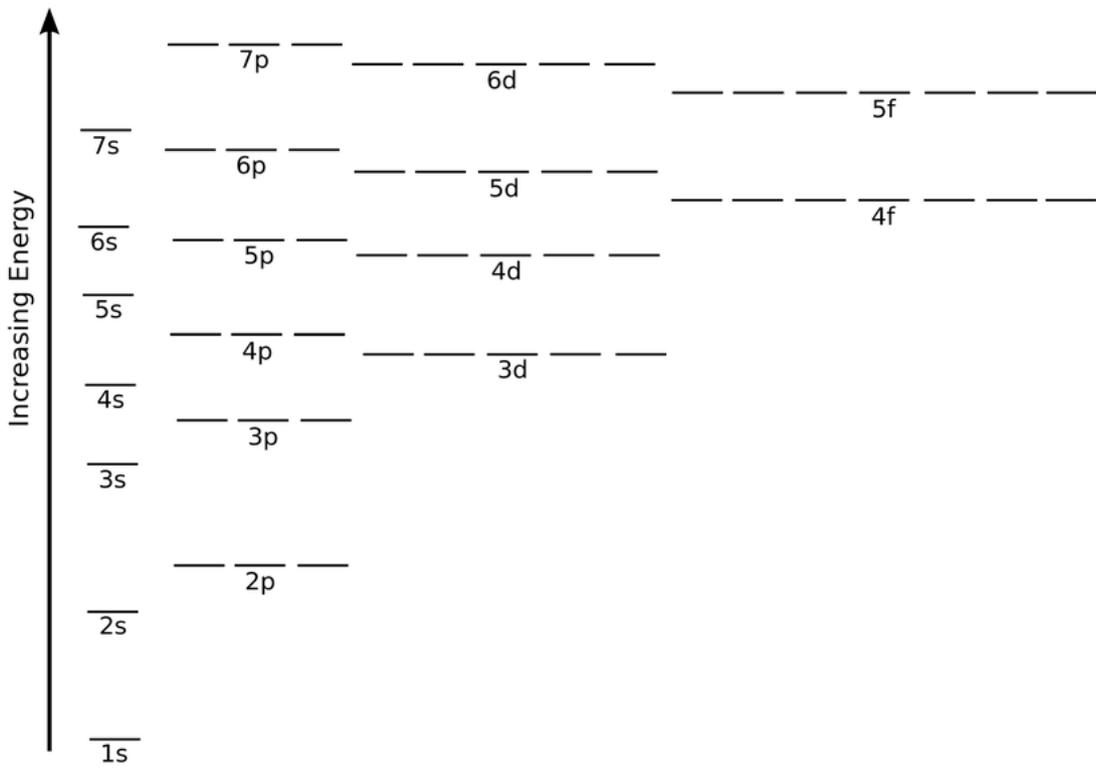
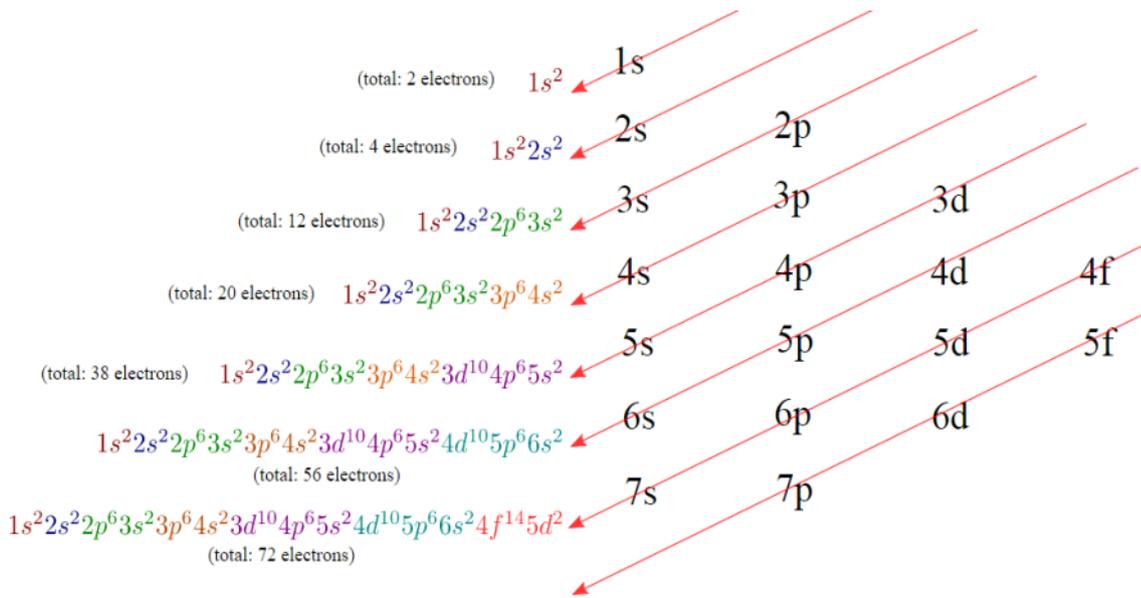


Notes:

Making Electron Configurations

- When filling atomic orbitals with electrons, we must fill from the lowest energy to the highest energy first
- This is known as the aufbau principle which states the aforementioned point
- When filling electron orbitals, we must also follow another rule known as Hund's rule which states that in a particular set of orbitals of the same energy, the lowest energy configuration for an atom is the one with the maximum number of unpaired electrons allowed by the Pauli exclusions principle
- Unpaired electrons are represented as having parallel spins
- In other words, as orbitals fill up, it is energetically favourable to half fill orbitals with electrons of the same spin and only when there is no more space should electrons be paired up





Unexpected Electron Configurations and Multivalent Charges

Formulas & Theorems Covered Today:



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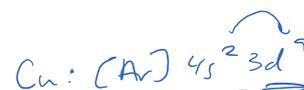


Notes:

Unexpected Electron Configurations

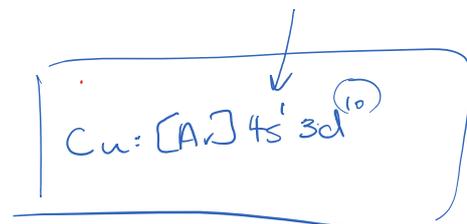
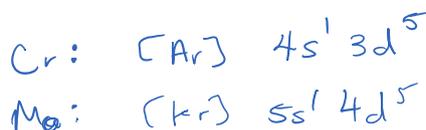
- During your last practice, you may have noticed that some elements do not follow the aufbau principle when it comes to their electron configurations

o Ex. Cr



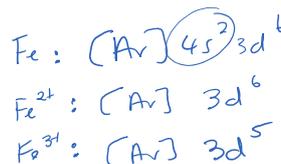
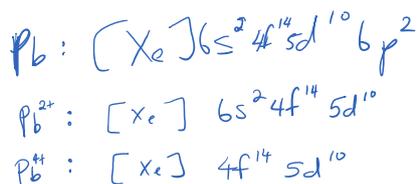
- This occurs because half-filled or fully-filled subshells are more stable than empty or partially filled subshells
- In the case of chromium, one of the 4s electrons will jump to fill the empty 3d orbital

o Ex. Cr



Explaining Multivalent Charges

- Some metals have multiple valences which can be explained using the rules that we learned about filling orbitals and the electron configurations
- As an example, Lead can have a charge of +2 or +4
- When we see Lead's electron configuration, we see it as $[\text{Xe}]6s^2 4f^{14} 5d^{10} 6p^2$
- You have also learned from your previous courses that the positive charges that occur in metal ions are caused by the loss of electrons
- When those electrons are removed, they are removed furthest from the nucleus to closest in that order
- In the example we see above with Lead, the furthest electrons from the nucleus are the two 6p electrons and next closest will be the two 6s electrons
- In order to achieve the +2 ion Lead will lose the two 6p electrons first, and then to achieve the +4 ion, it will lose in addition to the two 6p electrons, the two 6s electrons



Periodic Trends

Formulas & Theorems Covered Today:

- ★ Periodic Trends
- ★ Ionization Energy
- ★ Atomic Radius

Homework:

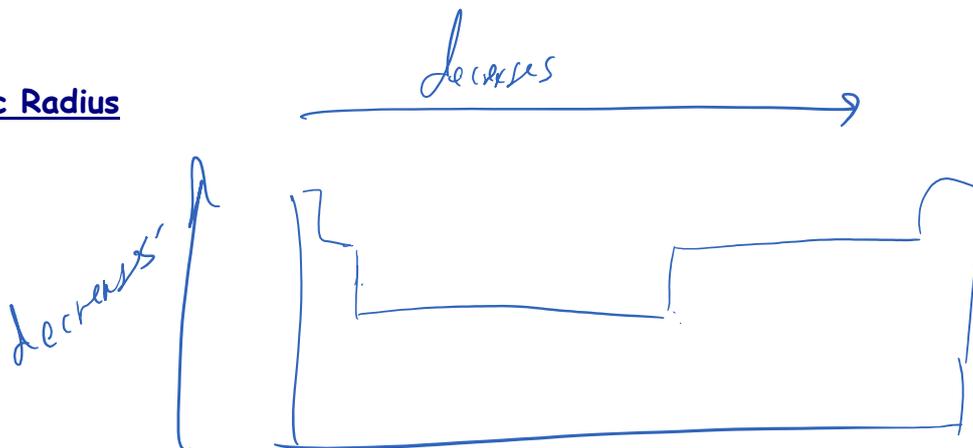
- 🏠 Paramagnetism Lab

Notes:

Periodic Trends

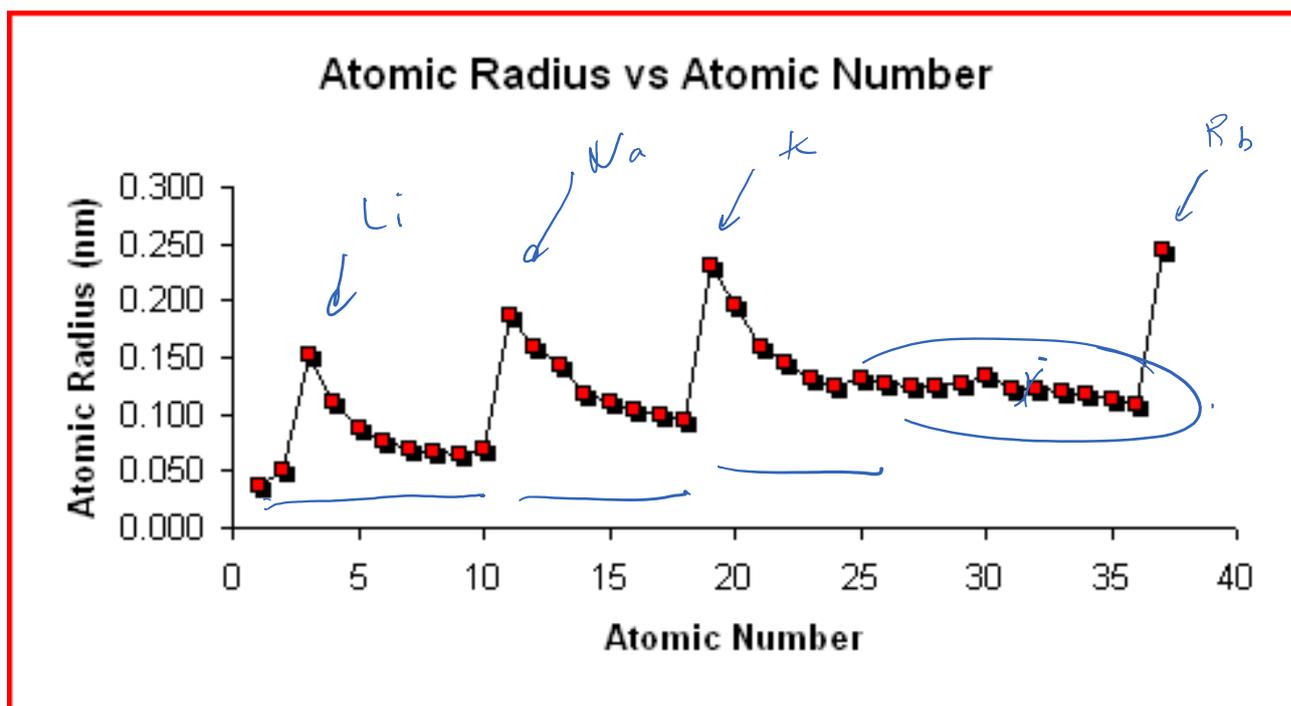
- As we have learned recently, atoms are not solid spheres
- Rather their volumes can be defined as probabilities of where electrons could be
- The size of an atom is dictated by many factors and is referred to as the atomic radius
- Scientists have determined the radius of an atom by looking at the distance between bonded nuclei
- In metals, as an example, the radius is half the distance between neighboring nuclei in a crystal of that metal element

Atomic Radius



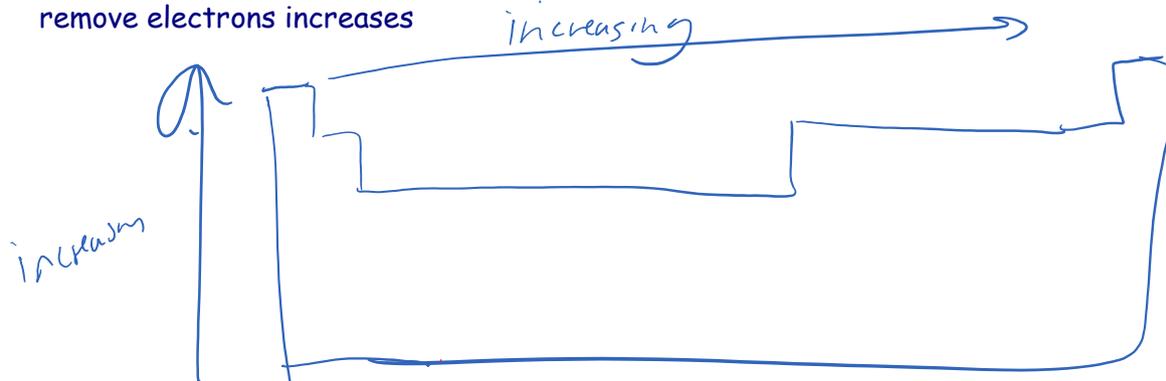
- All atomic radii are measured in picometers ($1 \times 10^{-12} \text{m}$)
- Understand that the nucleus of an atom is 250,000 times smaller than the atom itself, that is, if a nucleus were the size of a basketball, the closest electrons would be in Buffalo, NY
- There are two factors that affect the atomic radius
- The first is the change in "n" ←
- As you increase the energy level, the size of the orbital changes and the probability of finding the electrons further from the nucleus also changes
- This effectively makes the radius of an atom larger as you go down a column
- The second factor that affects the atomic radius is the nuclear charge
- As you move across a period on the table, the overall charge of the nucleus increases (you are adding protons)
- The protons have an electrostatic attraction to the electrons in the orbitals
- The reason that the atom's radius decreases as you move left to right on the periodic table is

- because the nucleus, now having more and more charge, pulls the electrons closer to it
- You may wonder and argue that at the same time that you are adding protons, you are also adding electrons and so the atoms radius should be getting bigger
 - That is true, but as long as you stay in the same period, you are staying in the same energy level and this the electrons having the same "n" are restricted to how far away they can be from the nucleus
 - This explains the trend as you proceed across a period
 - The largest change occurs in the elements found in the "s" and "p" orbitals
 - Transition metals change very little because they are part of the "d" orbitals and you will recall that the "d" orbitals are always one "n" less than the "s" and "p" orbitals

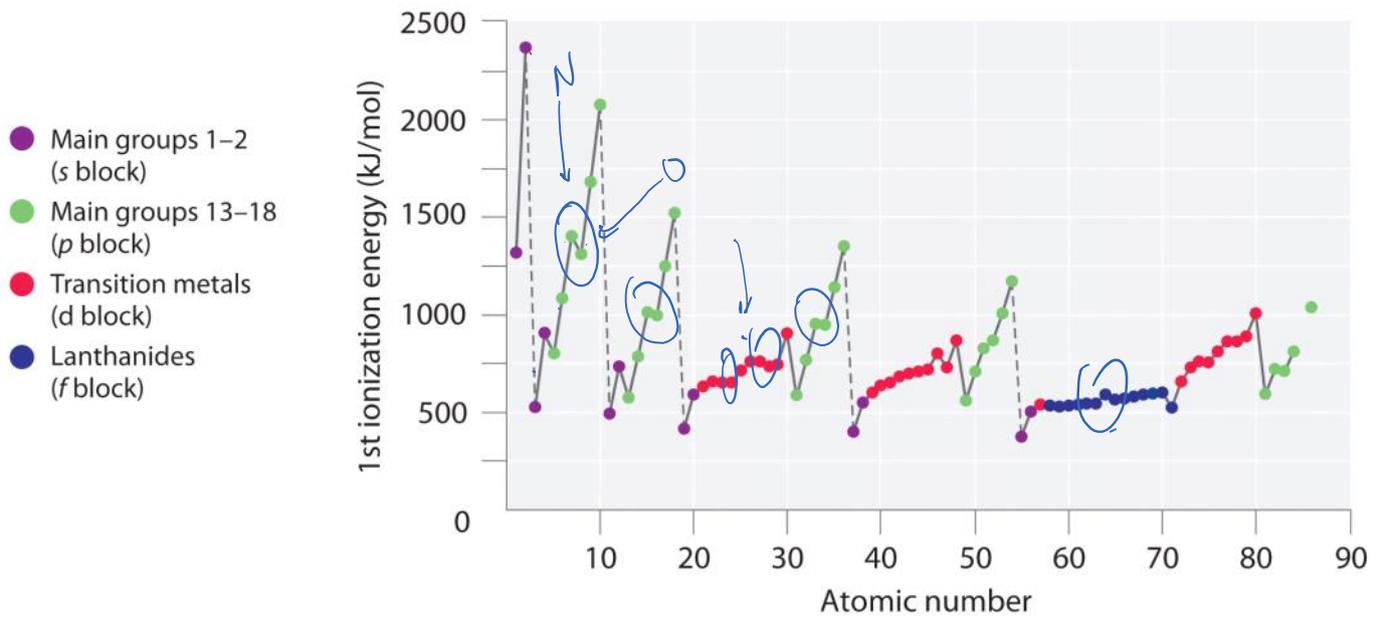


Ionization Energy

- Ionization energy is the amount of energy in $\frac{kJ}{mol}$ that is required to remove an electron from an atom
- Effectively, you are creating an ion, hence the name of the trend
- Every element has a 1st, 2nd, 3rd, nth, ionization energy
- As you remove electrons it becomes more and more difficult to remove more electrons so you will notice that as you move from 1st to 2nd and so on, that the amount of energy required to remove electrons increases



- Notice the trend is the opposite to the trend for the atomic radius
- The closer an electron is to the nucleus, the more difficult it is to remove it



Lewis Structures and Polarity

Formulas & Theorems Covered Today:
 ★ Lewis Structure

Homework:
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Notes:

Lewis Structures

- As we mentioned yesterday, Lewis structures are used to help determine how molecules are put together
- They can also be used to help us determine whether a molecule is polar, or non-polar
- The polarity of molecules, as you will see in the next few classes, will play a large role on the intermolecular forces that are so important for determining the physical properties that compounds possess

Lewis Structure Calculations

Step 1

- Place the least electronegative element at the center of the Lewis structure
- Write the symbols for the other atoms around the central atom
- Always place hydrogen and fluorine at end positions

Step 2

- Determine the total number of valence electrons present in the molecule or ion
- Be sure to account for charges on any ions
 - If a cation, remove that number of valence electrons
 - If an anion, add that number of valence electrons

Step 3

- Determine the total number of valence electrons needed to complete all atoms octets

Step 4

- Subtract the total number of valence electrons that exist in the atom or ion from the total number that you require to complete the octet (Step 3 - Step 2)
- Take that number and divide by 2 in order to determine the number of bonds (this is done because bonds are created by 2 electrons)

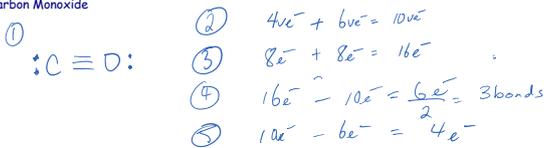
Step 5

- Subtract the number of shared electrons (bonding electrons) from the number of valence electrons (Step 2) to determine the number of non-bonding electron pairs (lone pairs)

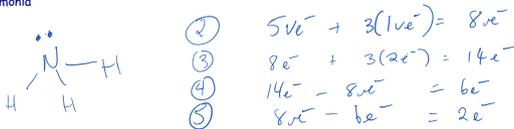
Step 6

- Place non-bonding pairs on outer atoms to fill their octets
- Any extra electrons are to be placed on the central atoms

Ex1. Carbon Monoxide



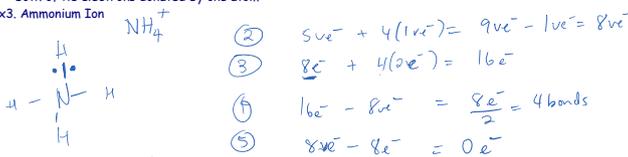
Ex2. Ammonia



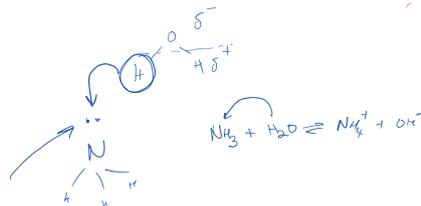
Co-ordinate Covalent Compounds

- Co-ordinate covalent compounds exist when a covalent bond that is found in a molecule has both of its electrons donated by one atom

Ex3. Ammonium Ion



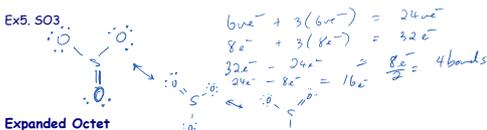
Ex4. Hydronium Ion H_3O^+



Resonance Structures

- These are compounds in which there is more than one Lewis structure possible
- When we measure these structures in a lab we do not see all of the possible structures but rather we see an average of all the bonds present

Ex5. SO2



Expanded Octet

- As we move up the periodic table, we add more and more energy levels and subsequently we add more orbitals
- Larger atoms have access to these orbitals and are able to add more than an octet full of electrons
- To create the Lewis structures for these elements, you follow a different set of rules below
- If you try and use the other rules, you will not produce enough bonds for the elements

involved

Step 1

- Place the least electronegative element in the middle and place all other elements around them with single bonds

Step 2

- Determine the number of valence electrons present in the atoms involved
- Make sure to account for the charges on ions as with the other rules

Step 3

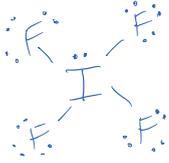
- Determine the number of non-bonding pairs but subtracting the number of bonding electrons (2xnumber of single bonds) from the number of valence electrons

Step 4

- Place the non-bonding electrons on the outer atoms first and any extras onto the central atom

Ex6. Phosphorus pentachloride

Ex7. IF_4^-



$$\textcircled{2} \quad 7e^- + 4(7e^-) - 1e^- = 34e^-$$

$$\textcircled{3} \quad 8e^- + 4(6e^-) = 40e^-$$

$$\textcircled{4} \quad 40e^- - 34e^- = \frac{6e^-}{2} = 3 \text{ bonds}$$

$$34e^- - 8e^- = 26e^-$$

VSEPR Theory

Formulas & Theorems Covered Today:



Homework:



Notes:

VSEPR Theory - Valence Shell Electron Pair Repulsion Theory

- This is a method that is used to distinguish the 3D conformation of a molecule in space
- We can no longer think of molecules or compounds as 2D structures drawn in our notebooks but must think of them as 3D structures as their 3D structure will play a large role in the intermolecular forces that are present in a molecule
- VSEPR is the theory that bonding pairs and lone pairs of electrons repel one another and that repulsion changes the shape that that molecules has in space
- Electrons are separated into orbitals in order to minimize the repulsive forces within atoms
- A lone pair of electrons will spread over a large area than a bonding pair as they do not have nuclei locking them into place
- The repulsion is said to be greatest between LP-LP, second greatest is between LP-BP and the least repulsion is between BP-BP
- Below are the 5 major electron group arrangements that can be found in molecules

Electron Groups	2	3	4	5	6
Geometry	Linear	Trigonal planar	Tetrahedral	Trigonal bipyramidal	Octahedral
Predicted Bond Angles	180°	120°	109.5°	90°, 120°	90°

Molecular Geometry

- Compare the Lewis structures for methane, ammonia, and water
- Ex1.

- Each of these has a tetrahedral confirmation as they each have 4 electron groups around the central atoms
- What we do see however, despite them having the same number of electron groups, is the effect that the type of electron groups has on the 3D shape of the molecules

VSEPR Notation

- The VSEPR notation, used in conjunction with the Lewis structure and the VSEPR shapes chart

VSEPR Notation

- The VSEPR notation, used in conjunction with the Lewis structure and the VSEPR shapes chart will help us determine the 3D shape that a molecule will have

A - Represents the central atom

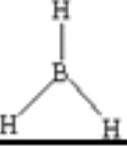
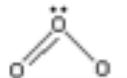
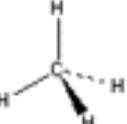
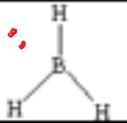
X - Represents the bonding pairs

E - Represents the non-bonding pairs

H ₂ O	AX ₂ E ₂	Angular
NH ₃	AX ₃ E	Trigonal Planar
CH ₄	AX ₄	Tetrahedral

Ex2. Sulfur Trioxide

Ex3. XeF₄

Pairs of Electrons	Pairs of Bonding Electrons	Pairs of Lone Electrons	Electron Distribution	Molecular Geomtry	Bond Angle	Diagram
2	2	0	linear	linear	180	
3	3	0	trigonal planar	trigonal planar	120	
	2	1		bent	117	
4	4	0	tetrahedral	tetrahedral	109.5	
	3	1		trigonal pyramidal	107	
	2	2		bent	104°	

				pyramidal		
	2	2		bent	104°	
5	5	0	trigonal bipyramidal	trigonal bipyramidal	120 + 90	
	4	1		seesaw	117 + 90	
	3	2		T-shaped	90	
	2	3		linear	180	
6	6	0	octahedral	octahedral	90	
	5	1		square pyramidal	87	
	4	2		square planar	90	

Chemical Bonding

Formulas & Theorems Covered Today:

- ★ Ionic Bonding
- ★ Covalent Bonding
- ★ Metallic Bonding

Homework:



Notes:

Chemical Bonding

- Of all the elements that occur naturally or synthetically, it is only the noble gases that exist in uncombined atoms
- The rest are held together in molecules or compounds by chemical bonds
- The reason that bonding occurs in the first places is because of energy
- In nature, lower energy systems are always favoured over higher energy systems
- The reason for this favourability is that the lower energy systems are more favourable

Simple Lewis Structures

- For the bonding to take place, only the valence electrons take part in the bonding
- In grade 10, you would have learned how to make Lewis structures for elements and in grade 11, you would have learned about Lewis structures for ionic compounds or simple covalent compounds

- Ex1. Elements

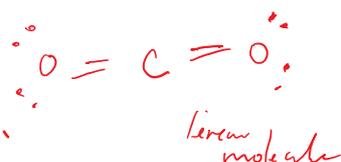


- Ex2. Ionic/Simple covalent compounds



- In this course, we will use Lewis structures to represent bonding of more complex compounds
- When we would like to represent a bond, we will use a line, and when we would like to represent lone pairs of electrons, we will use a pair of dots

- Ex3. Carbon Dioxide

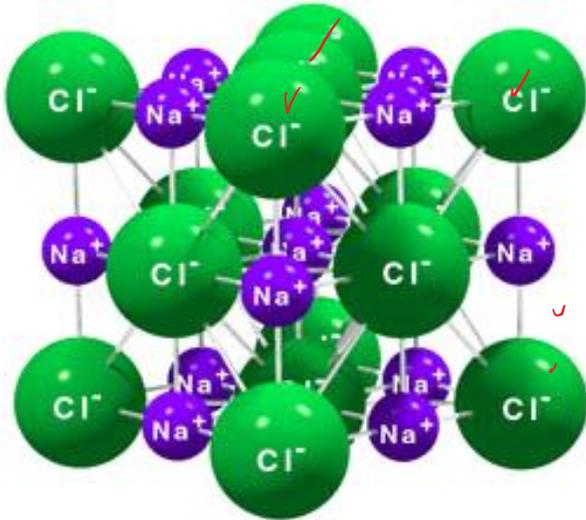


$$\begin{aligned} 4ve^- + 2(6ve^-) &= 16ve^- \\ 8e^- + 2(8e^-) &= 24e^- \\ 24e^- - 16ve^- &= 8e^- = \frac{8}{2} = 4 \text{ bonds} \\ 16ve^- - 8e^- &= 8e^- \end{aligned}$$

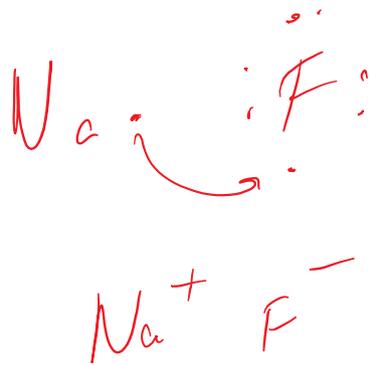
Ionic Bonding

- Ionic bonding is the force of attraction between oppositely charged ions (cations and anions)
- This occurs between elements with large differences in electronegativity (you learned this last year)
- Typically we see a metal with a low electronegativity bond with a non-metal with a high electronegativity
- These compounds are usually arranged in a crystal lattice array and are very strong

difficult to separate by simply heating

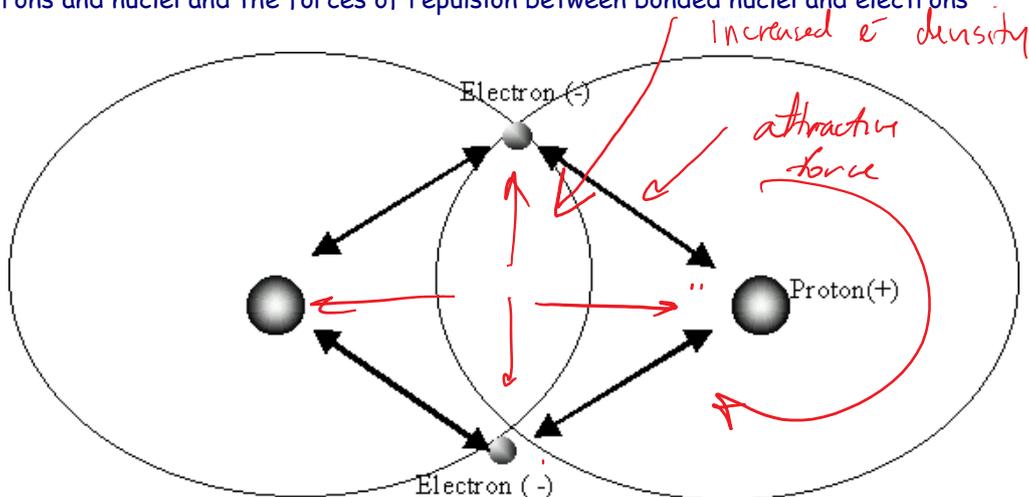


- The above image is an example crystal lattice of sodium chloride (or table salt)
- This type of bonding boasts electrons that have been transferred from an "s" orbital (metal) to the "p" (non-metal) orbitals
- Another way to represent this is to look at the Lewis structures or to represent the bonding using orbital diagrams
 - Ex4. Sodium Fluoride

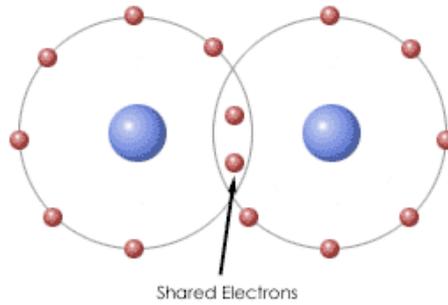


Covalent Bonding

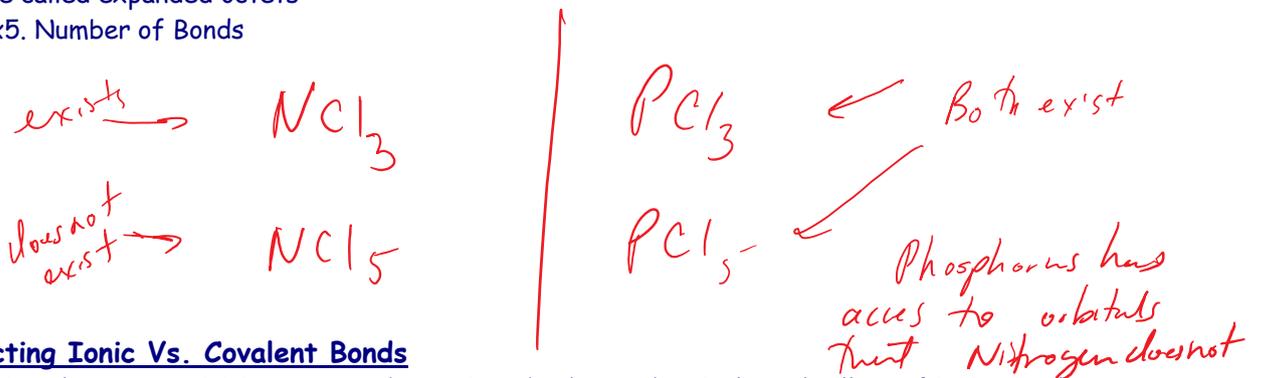
- This type of bonding involves a delicate balance between the forces of attraction between electrons and nuclei and the forces of repulsion between bonded nuclei and electrons



- If we look at the word covalent, it simply means to share the valence shell
- The reason that this even happens is so that the atoms involved in the covalent bonding achieve noble gas electron configuration (full electron shell)
- A single shared pair of electrons fills the valence level for both atoms

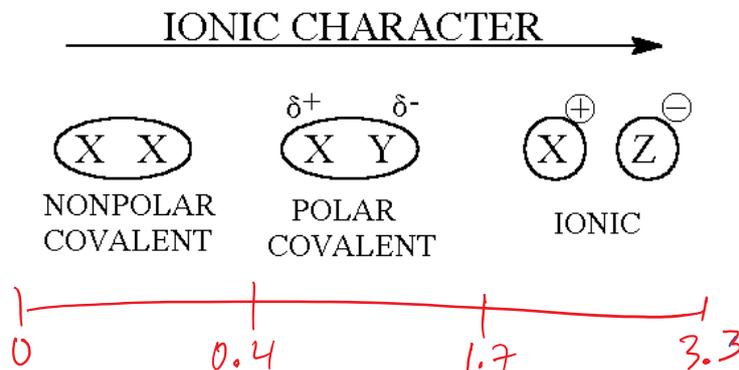


- Covalent bonding at the 2P level must always obey the octet rule
- There are however examples of compounds that can break the octet rule and these examples are called expanded octets
- Ex5. Number of Bonds



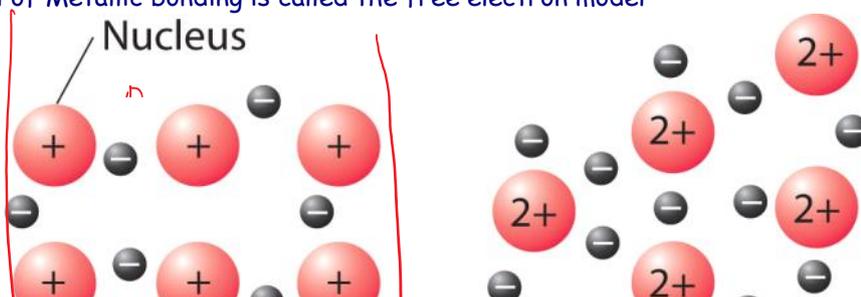
Predicting Ionic Vs. Covalent Bonds

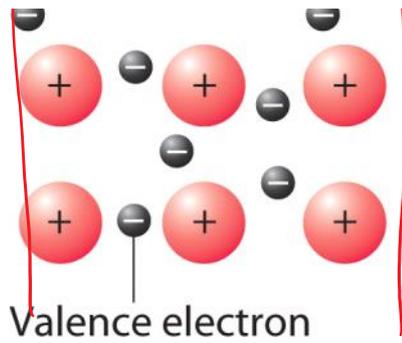
- Using electronegativities you can determine whether a chemical bond will be of ionic, covalent, or polar covalent nature
- As you have learned in previous classes, electronegativity is the "want" or "pull" that an element has on electrons and in bonding it is on its neighbours electrons
- Below you will see a bonding continuum that will help determine where a particular type of bond falls



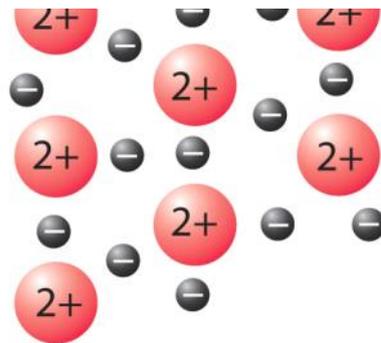
Metallic Bonding

- Metals are generally conductive, malleable, ductile and solid
- Metals, and metal bonding is different than other types of bonding in that they share electrons throughout their structure (delocalized electrons)
- This model of Metallic bonding is called the free electron model





(a) Group 1 metal



(b) Group 2 metal

- Metallic bonds are formed between the cations and the "pool" or "sea" of valence electrons
- This explains why metals are good conductors and can also explain why they are so easily deformed (malleable)

Bond Vs Molecular Polarity

Formulas & Theorems Covered Today:

- ★ Bond vs Molecular Polarity
- ★
- ★

Homework:



Notes:

- The shape and type of bonds in a molecule will dictate the polarity of that molecule
- This will then dictate the physical properties as the molecular polarity will tell us which type of intermolecular forces are present in the compound
- As we saw in the chemical bonding note, the bonding continuum is used to determine polarity of the bonds
- The VSEPR shape along with the bond polarity will dictate the molecular polarity

$$\Delta EN = \text{Bond polarity}$$

2 conditions you must satisfy for a molecule to be polar

1. you must have polar bonds
2. molecule must be asymmetrical



$$\Delta EN_{C-H} = 2.55 - 2.2 = 0.75$$

Pure covalent bond

Non-polar molecule

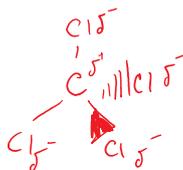
Chloromethane



polar bonds present ✓
asymmetrical molecule ✓

polar molecule

Carbon tetrachloride

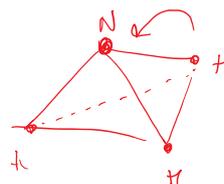
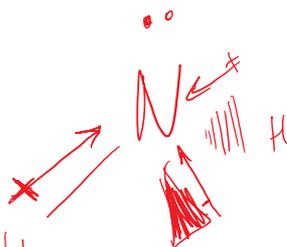


$$\Delta EN_{C-Cl} = 3.16 - 2.55 = 0.61$$

Polar covalent bond

Non-polar molecule

=)





Intermolecular Forces

Formulas & Theorems Covered Today:

- ★ 5 Types of Intermolecular Forces
- ★
- ★

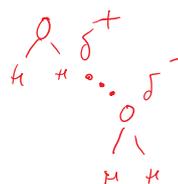
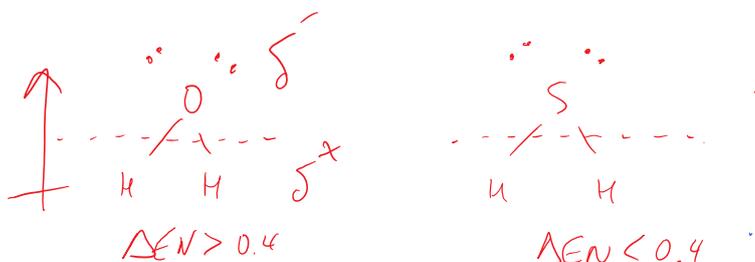
Homework:



Notes:

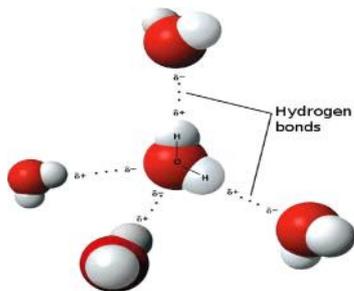
Intermolecular Forces

- Water and hydrogen sulfide are two molecules that are of similar shape and both have an angular shape
- Interestingly, these two compounds, despite them being roughly the same size, and the same shape have very different physical properties
- Water boils, under normal pressure conditions at 100°C while hydrogen sulfide boils at -61°C
- So why the difference?
- Everything that we have learned about thus far in chemistry has been about intramolecular forces, that is forces that hold molecules together
- These are the forces that we know as chemical bonds and are responsible for the chemical properties of that particular element
- The physical properties are influenced by the intermolecular forces (between molecules)
- These were extensively studied by a scientist named Johannes van der Waals
- Intermolecular forces are often times called van der Waals forces
- They can be categorized into 5 major groups
 - Dipole-Dipole forces
 - Ion-Dipole forces
 - Induced-Dipole forces
 - Hydrogen Bonding
 - London Dispersion forces
- Why the difference?



Dipole-Dipole Forces

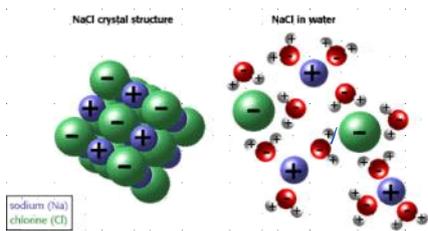
- In the liquid state, polar molecules, will align their oppositely charged polar regions to one another
- The attraction by these polar regions is what we call the dipole-dipole force
 - Ex1.



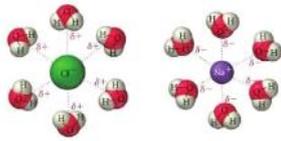
- As a result, polar molecules will tend to be liquid at room temperature, while other, similarly sized non-polar molecules, will be gasses at room temperature

Ion-Dipole Forces

- The reason that salt, which is an ionic solid, can dissolve in water is because of the ion-dipole interaction that exists between the water molecules and the ions present in the salt crystal
- An ion, like sodium or chlorine, will be attracted to the partially negative or partially positive ends of the water molecule respectively
- The reason that this happens is because the forces that are exerted by the ions on the water molecules, and vice versa, are enough to overcome the ion-ion interaction in the salt crystals
 - Ex2.



o Ex3.

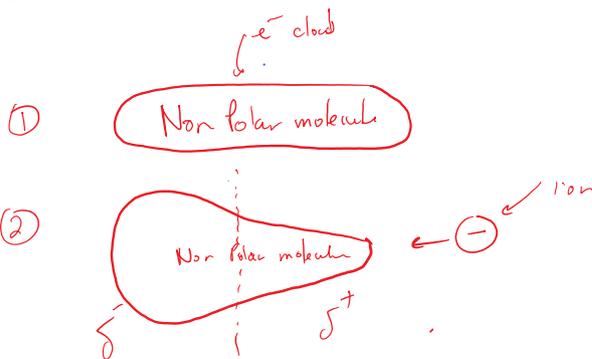


Induced-Dipole Forces

- When an object is charged by induction, a charge is forced on an object when the electrons within that object move to one side of that object, pushed by some source of charge
- In chemistry, the charge that can push these electrons is either an ion or another polar molecule

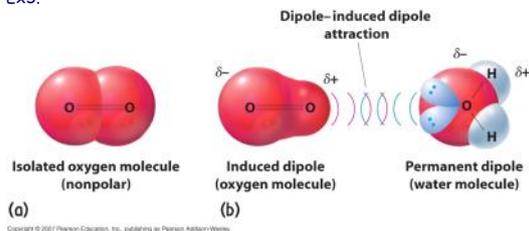
o **Ion-Induced Dipole**

- This happens when an ion gets close enough to a molecule that it distorts the electron cloud and either pulls or pushes electrons based on its charge
- Ex4.



o **Dipole-Induced Dipole Forces**

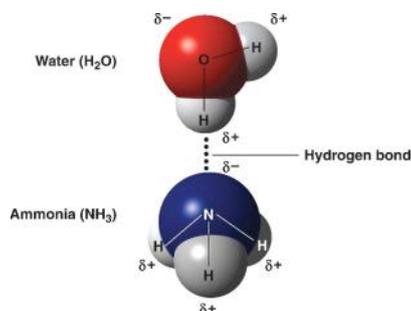
- These follow the same principle as the ion-induced dipole forces except instead of an ion inducing the molecule, it is another polar molecule
- Ex5.



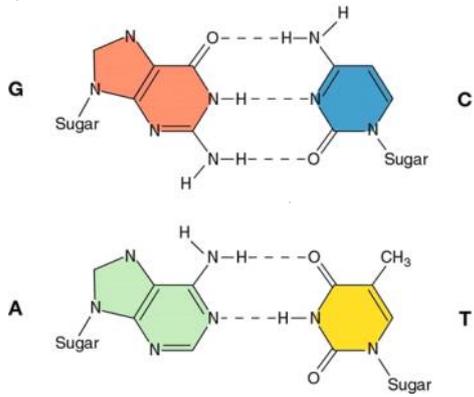
Hydrogen Bonding

- This occurs when an element within a molecule becomes trapped between an electronegative element that it is bonded with and an electronegative element in a neighbouring molecule
- This type of dipole-dipole interaction is the strongest of all types of intermolecular forces but is not a bond, as the name would suggest
- This type of intermolecular force is responsible for life on Earth, as it is the dominant force in liquid water and it is what makes ice less dense than water
- It is also the force that is responsible for holding DNA together

o Ex6.



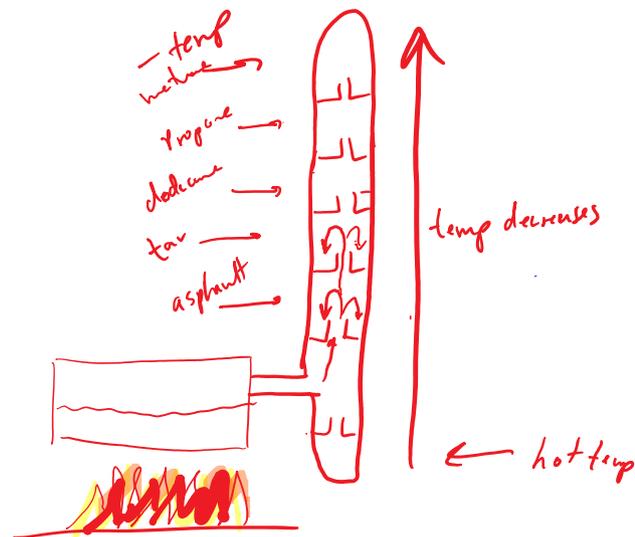
Ex7.



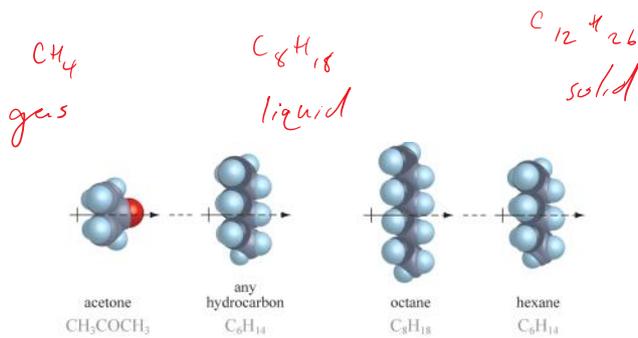
London Dispersion Forces

- This applies to all molecules but is of particular importance and interest to non-polar molecules as it is the dominant force present
- Bond pairs of electrons are in constant motion between the nuclei
- This motion can cause what is called a momentary uneven distribution of the electrons and creates a momentary dipole
- This creates a situation where one non-polar molecule becomes momentarily polar
- At that instant in time, that molecule may be able to create some level of attraction to its neighbouring non-polar molecule, either through inducing a dipole or attracting an oppositely charged dipole
- The larger a molecule is, the greater the total number of bonds there are within the molecule, which gives a greater opportunity for these momentary dipoles to exist causing a greater London dispersion forces present
 - Ex8.

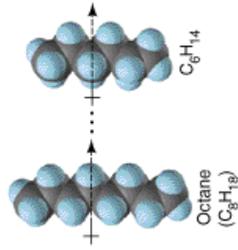
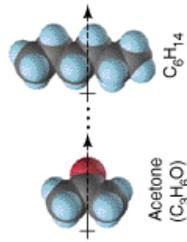
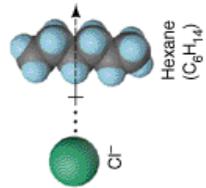
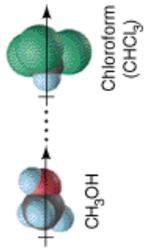
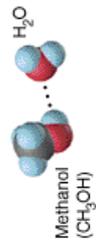
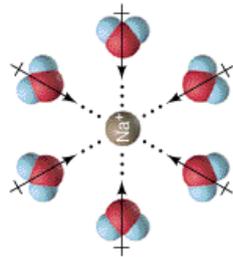
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Methane	Octane	Dodecane
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Summary



Solids and their Properties

Formulas & Theorems Covered Today:



Homework:



Notes:

