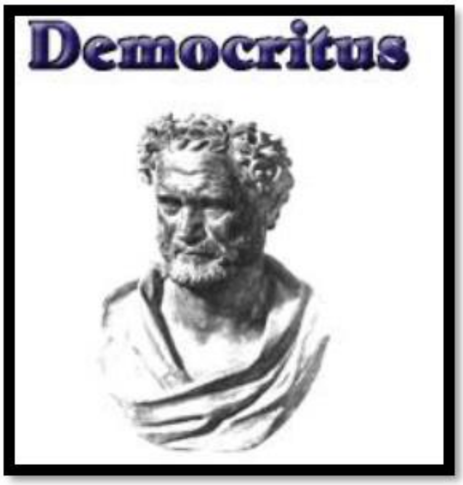
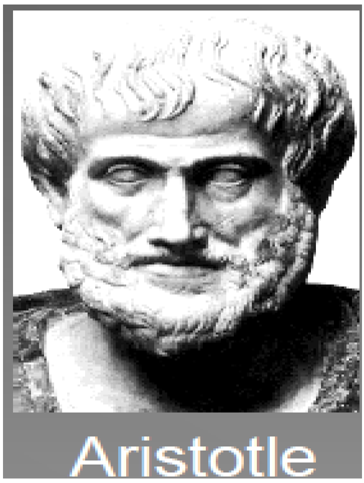
History of Atomic Theory



Timeline: 400 BC Scientist: Democritus (Greek Philosopher) Democritus was a Greek philosopher who was the first person to use the term **atom** (atomos: meaning indivisible). He thought that if you take a piece of matter and divide it and continue to divide it you will eventually come to a point where you could not divide it any more. This **fundamental** or **basic** unit was what Democritus called an atom.



Timeline: 350 B.C-Aristotle modified an earlier theory that matter was made of four “elements”: earth, fire, water, air.

•Aristotle was wrong. However, his theory persisted for 2000 years.

Timeline: 1800’s Scientist: John Dalton was the first to adapt Democritus’ theory into the first modern atomic model.

JOHN DALTON'S ATOMIC MODEL

1.Atoms were solid spheres.

2.All substances are made of atoms; atoms are small particles that cannot be, divided or destroyed.

3.All atoms of an element are identical in size, mass, shape and other physical properties.

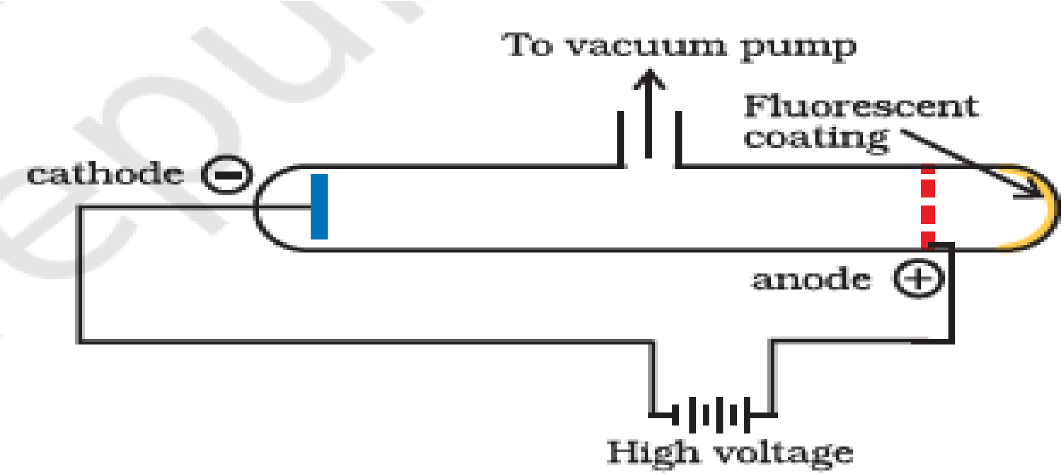
4.Atoms of different elements have different in size, mass, shape weights and different chemical properties.

5.Atoms combine to form compounds

Cathode Ray Discharge Tubes

Timeline: 1890's Scientist: J.J Thomson was a physicist who is credited for discovering the electron. He used his research on cathode ray tube technology in this discovery. Cathode ray tube is made of glass containing two thin pieces of metal, called electrodes, sealed in it. The electrical discharge through the gases

could be observed only at very low pressures and at very high voltages. The pressure of different gases could be adjusted by evacuation. When sufficiently high voltage is applied across the electrodes, current starts flowing through a stream of particles moving in the tube from the negative electrode (cathode) to the positive electrode (anode). These were called **cathode rays or cathode ray** particles.



The flow of current from cathode to anode was further checked by making a hole in the anode and coating the tube behind anode with phosphorescent material zinc sulphide. When these rays, after passing through anode, strike the zinc sulphide coating, a bright spot on the coating is developed.

The results of these experiments are summarized below.

(i) The cathode rays start from cathode and move towards the anode.

(ii) These rays themselves are not visible but their behavior can be observed with the help of certain kind of materials (fluorescent or phosphorescent) which glow when hit by them.

(iii) In the absence of electrical or magnetic field, these rays travel in straight lines

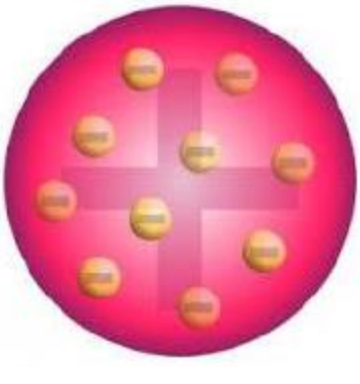
(iv) In the presence of electrical or magnetic field, the behavior of cathode rays are similar to that expected from negatively charged particles, suggesting that the cathode rays consist of negatively charged particles, called electrons.

(v) The characteristics of cathode rays (electrons) do not depend upon the material of electrodes and the nature of the gas present in the cathode ray tube.

Thus, we can conclude that electrons are basic constituent of all the atoms.

THOMSON'S ATOMIC MODEL

Proved that an atom can be divided into smaller parts. And proposed that an atom possesses a spherical shape (radius approximately 10^{-10}m) in which the



positive charge is uniformly distributed. The electrons are embedded into it in such a manner as to give the most stable electrostatic arrangement. Many different names are given to this model, for example, plum pudding, raisin pudding or watermelon.

Rutherford's Nuclear Model of Atom

Rutherford and his students bombarded very thin gold foil with a α-particles. A stream of high energy a α-particles from a radioactive source was directed at a thin foil (thickness ~ 100 nm) of gold metal. The thin gold foil had a circular fluorescent zinc sulphide screen around it.

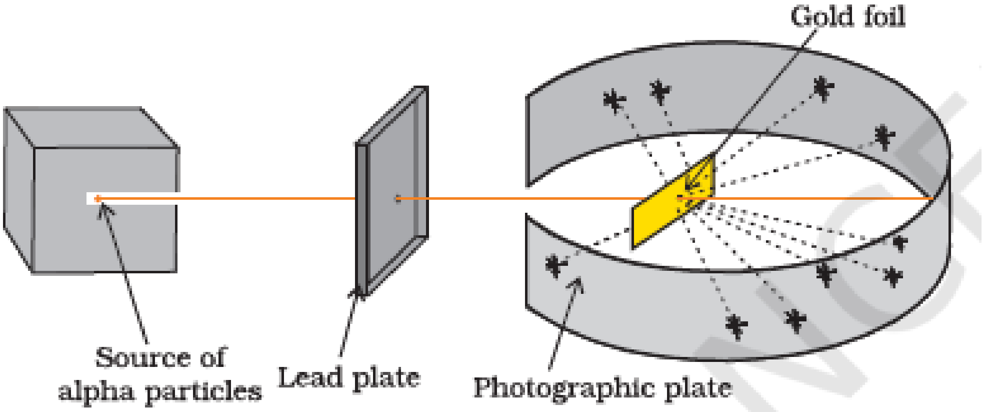
Whenever a α–particles struck the screen (Without putting a gold slide). a tiny flash of light was produced at that (A) point.

After putting the foil of gold it was found that :

i. most of the a α–particles passed through the gold foil undeflected

ii. Small fraction of the a α–particles was deflected by small angles.

iii. Very few a α–particles (~1 in 20,000) bounced back, that is, were deflected by nearly 180°.



A. Rutherford's scattering experiment

On the basis of the observations,

Rutherford drew the following conclusions

i. regarding the structure of atom :

ii. Most of the space in the atom is empty as most of the α– particles passed through the foil undeflected.

iii. A few positively charged a α–particles were reflected. The reflection must be due As a result of a colloide with a large body that can not be radiated penetrate.

iv. A few positively charged a α–particles were deflected. The deflection must be due to large repulsion force showing that the positive charge of the atom is not spread throughout the atom as Thomson had proved. The positive charge has to be concentrated in a very small volume that repelled and deflected the positively charged α–particles. This very small portion of the atom was called nucleus.

Rutherford's Nuclear Model of Atom

The Atom:

Although it has very small size but it has a complicated structure that resembles the solar system in which electrons revolve around the central nucleus in orbits as planets revolve around the sun.

2- The Nucleus:

Is much smaller than the atom. Located in the centre of the atom with positive charge. There is a big space between the nucleus and orbits of electrons, so most of the atom is a space. Most mass of the atom is concentrated in the nucleus and the mass of electrons is very small and can be neglected.

3- Electrons:

1-Have negligible mass compared to that of the nucleus.

2-Number. of electrons with negative charge are equals to the number of protons with positive charge so the atom is electrically neutral.

3- Electrons revolve around the nucleus in a fixed orbit

4- The electrons are affected by two forces equal in strength but in opposite direction, which are :

a- Attraction Force of the nucleus to electrons.

b- Centrifugal force due to velocity of electron around the nucleus.

Objection to the Rutherford Model:

According to Maxwell's theory: If a charged electron particle moves around another particle charged with a opposite charge, the electron loses part of its energy coming out in the form of radiation, thereby reducing the orbit gradually. And that is mean the electron finally fall in nucluse (and that never done).

Atomic spectra and its explanation (Bohr theory)

The study of atomic spectra is considered the key which solved the puzzle of atomic structure. That was the work of Bohr.

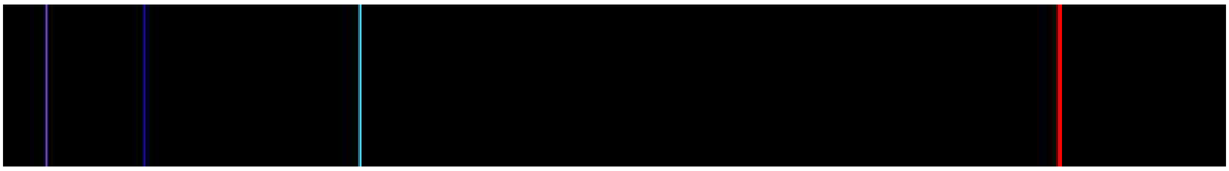
On heating atoms of a pure element in gaseous or vapour state to high temperature or exposing them to low pressure inside electric discharge tube, they emit radiation known as line spectrum. On examining this radiant light by a spectroscope, we observe a group of small number of restricted coloured lines separated by dark areas so it's called line spectrum. It was found experimentally that the spectral lines are essential characteristic for each element i.e. there are no two elements that have the same spectral lines.

Spectrum:- The visible part of the light that analyzed using an optical analyzer (S.S)

Atomic spectra are of two types:

1) Linear emission spectrum

Are separate colored lines present in dark areas radiated by the excited element's



2) Linear absorption spectrum

Black lines in the continuous spectrum

Bohr’s Atom

Bohr's model was based on his observations on the spectrum produced by heating the hydrogen atom. When heating this atom emits radiation, then we expose it to a triple prism. The triple prism makes the white light deviate, producing all the colors of the visible spectrum. Each color corresponds to a specific amount of energy; however, when light emitted from a hydrogen atom is passed through a triple prism, only certain colors of light are seen. That's why Bohr says that electrons have specific quantities of energy in an atom, and by the colors emitted by a hydrogen atom.

Bohr’s Model

Studying the line spectra of hydrogen atoms, Bohr was able to reach his atomic model

Bohr Postulates

Bohr adopted some of Rutherford's postulates about atomic structure;

1 -A positively charged nucleus exists in the center of the atom.

2- Atom is electrically neutral as number of protons equals to number of electron's.

3- Electrons orbit the nucleus only in a definite allowed energy levels and they cannot be found at intermediate distances.

4- Electrons revolve around the nucleus in orbits due to centrifugal and attraction forces.

5- Electrons orbit the nucleus in a rapid movement without gaining or losing energy.

6- Each electron in the atom has a definite amount of energy depending on the distance between its energy level and the nucleus; the energy of any level increases as its radius increases.

7- The maximum number of energy levels in atoms in their ground state (unexcited) is only seven (K, L, M, N, O, P, Q). Each level has energy expressed by a whole number called principle Quantium. Number from (1 to 7).

8-When atom is excited by heating (Quantum) or by electric discharge the electron will transfer to a higher Enrgy. level agrees with the absorbed quantum. The excited electron in the higher Energy. level is then unstable, so it returns to its original level losing the same quantum of energy, which it gained during excitation in the form of radiation have definite wavelength and frequency.producing a characteristic spectrum line

REMARK:

1-Quantum Is defined as the amount of energy gained or lost when an electron jumps from one Energy level to another.

2- The difference in energy between levels (Quantum) is not equal i.e. the difference in this energy decreases further from the nucleus.

3- The electron does not move from its level to another unless the energy absorbed or emitted is equal to the difference in energy between 2 levels i.e. one quantum

.

(*There is not half quantum for instance). Quantum can't be divided or doubled*

The success of the Bohr model

1- It explained hydrogen atom spectrum.

2- He introduced the idea of quantum number to detect energy of electrons in energy levels.

3- He proved that electrons during rotation around the nucleus in ground state do not radiate energy, so they will not fall back to the nucleus .

4- (a reconcilion between Rutherford and Maxwell

The Limitation of the Bohr model

1-Bohr failed to explain the spectrum of any other element even that of Helium except hydrogen .

2-He considered the electron as negative charged particle only and did not consider that it also has wave properties.

3-He postulated that it is possible to determine precisely both speed and location of an electron at the same time. This is experimentally impossible.

4-He described the electron when moving in a circular planer orbit, Later it was confirmed that hydrogen atom has 3 dimensional co –ordinates

Sommerfeld atomic model

An improved model of the Bohr atomic model was proposed by Summerfield in 1916. Summerfield agreed to all the hypotheses in Bohr's theory except for the circular orbits of electron motion. According to Bohr's theory, when an electron falls from a higher energy to a lower energy level, it radiates one spectral line, but using spectral devices with high analytical power, shows that each spectral line consists of at least two spectral lines separated by very little distance.

So Sommerfield made a modification of Bohr's theory based on:.

1- Each of the energy levels produced by Bohr's model includes a number of sub-levels equal to the level number so that one orbit is circular and the rest is oval.

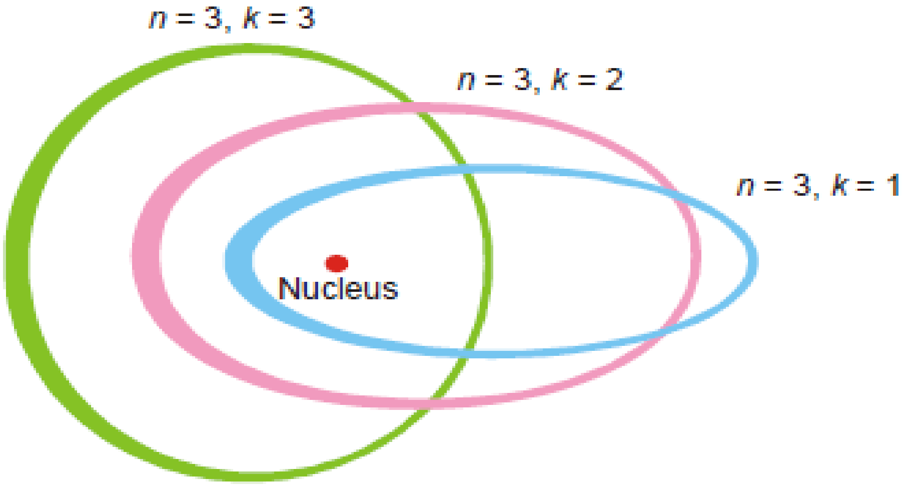
2- To determine the energy of the oval orbit, sommerfield entered a new quantum number known as the number of osmotic quantities (K), which is an integer that expresses the energy of the electron in the oval orbit (sub-level number). The oval in which the electron moves.

3- The speed of an electron's movement changes according to its position in the orbit so that it increases as it approaches the nucleus and decreases as it moves away from the nucleus.

4- Using n and K, Summerfield was able to determine the value between the axis length of the orbital using the simple equation(K=n-1)

$b/a=(K+1)/n$

a,b= the axis of the oval orbital



Modern Atomic Theory

1- The wave nature of the electron

All previously experimental considered the electron just a negatively charged particle but de Brawley assumed that the electron has a dual nature.

De Brawley considered that Every moving body (such as electron or the nucleus of an atom or whole molecule) is associated with (accompanied by) a wave motion (or matter waves) which has some properties of light waves.

a) It is a material particle b) It has wave properties

Electromagnetic Wave	Matter wave
They are separating from the moving body	They are not separating from the moving body
Their speed is equal to the speed of light	Their speed is not equal to the speed of light

plank Einstein

$$E = h\nu \longrightarrow E = mc^2 \longrightarrow$$

de Brawley

لذلك يكون

$$\longrightarrow mc^2 = h\nu = h\frac{c}{\lambda} \longrightarrow mc = \frac{h}{\lambda} \longrightarrow \boxed{\lambda = \frac{h}{mc}}$$

2- The principle of uncertainty

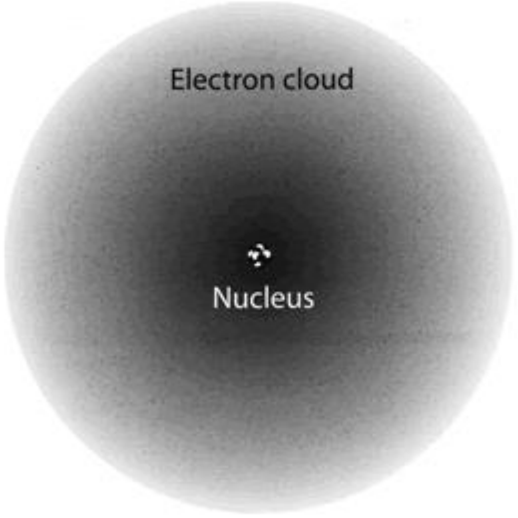
According to Heisenberg's uncertainty rule, it is impossible to know the velocity and the location of the electron around the nucleus at the same time.

But, it is correct is to say that the electron may be somewhere around the nucleus as much as possible. That is called the electron clouds.

3- Quantum Numbers (Schrödinger equation)

In 1926, the Austrian scientist, Schrodinger applying the ideas of Planck, Einstein, De Broglie, and Heisenberg established the wave mechanical theory of the atom and managed to derive a wave equation that could describe the electron wave motion in the atom. **On solving** Schrodinger's equation:.

It is possible to determine the allowed energy levels and to define the region of space around the nucleus where it is most probable to find the electron in each energy level. As a result of Schrodinger's work, our concept of the electronic motion around the nucleus has changed. Instead of speaking about the stable circular "orbits" of particular radii, and the areas between these orbits as being completely forbidden for electrons. The concept of electron cloud used to express the region of space around the nucleus where the possibility of finding the election in all distances, and directions. Inside the electron cloud there are areas that have a great possibility of finding an electron in it called orbital.



REMARK

1-Electron cloud Is defined as the region of space around the nucleus where the possibility of finding the election in all distances, and directions

2- Orbital :The area of space inside the electron cloud where there is a great probability for finding electrons

The mathematical solution of the Schrodinger equation introduced four numbers which are called quantum numbers.

Quantum Numbers:

They define the energy, shape, number and direction of orbitals

1- Principle Q.no (n).

1-The principal quantum number (n) describe the distance of the electron from the nucleus

2-Order of principle energy levels their number in the heaviest known atom in the ground state is seven

2-Number of electrons required to fill a given energy. level = two times the square of the level no ($2n^2$).

-1st E.L	**K**	**Is filled with**	**2 electrons**
-2nd E.L	**L**	**Is filled with**	**8 electrons**
-3rd E.L	**M**	**Is filled with**	**18 electrons**
-4th E.L	**N**	**Is filled with**	**32 electrons**

2- Subsidiary Q.no (L).

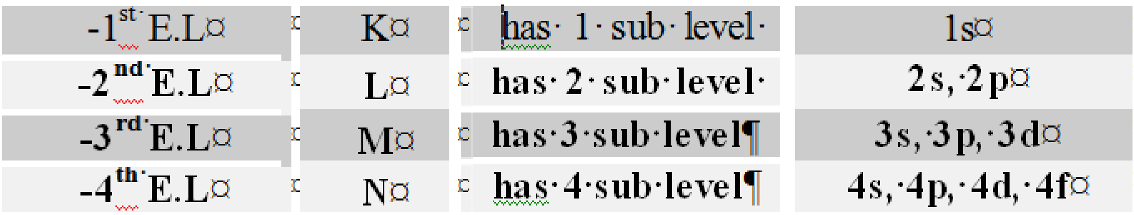
1- Used to detect the number of sub levels in each energy. level.

2- The energy sub levels take the symbols s, p, d, f. this is shown by the scientist Somerfield. When he used a spectroscope which has a high resolving power, he found that the single line (which represents electron transition between two different energy levels) is indeed a number of fine spectral lines which represents electron transition between very near energy levels (sublevels).

3-Number of sublevels in each energy level = order of principle energy level (n).

4-Energy of sub levels of same Energy level is not equal.

5- Values of subsidiary quantaum (L) =(n-1).



-1st E.L	K	has 1 sub level	1s
-2nd E.L	L	has 2 sub level	2s, 2p
-3rd E.L	M	has 3 sub level	3s, 3p, 3d
-4th E.L	N	has 4 sub level	4s, 4p, 4d, 4f

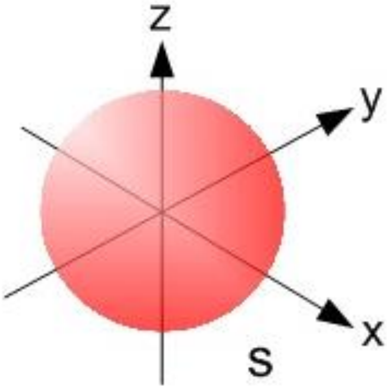
S=0 P=1 d=2 F=3

f>d>p>S

6-Energy of sub levels of same energy. level is not equal

3-Magnetic Q number (m)

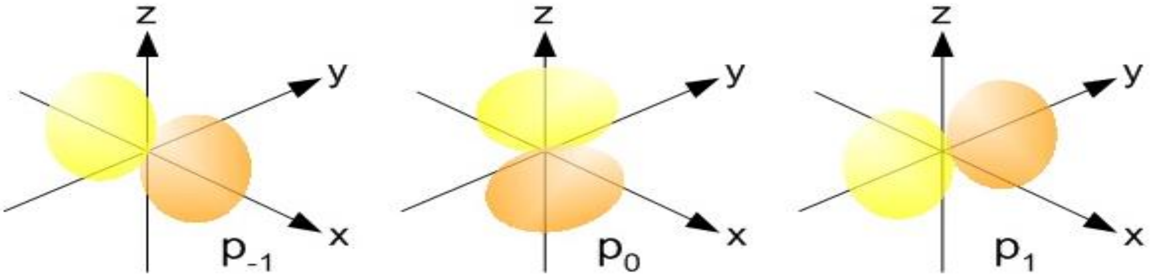
1- Used to detect no of orbitals in each energy sub level and their direction in space. which equal to (m=2l +1) where (l) is the value of the number of Subsidiary quantum, and the value of the number of magnetic quantum range between (-l) and (+ l).



2- Sublevel (S) **(m = 2 x 0 +1 = 1)** So the level (s) has one direction in the space and as such it has a spherical shape around the nucleus

3- Sublevel (P) has 3 orbitals **(m = 2 x 1+1 = 3)**.

Therefore, the level (p) has three directions in the space (Px, Py, Pz) is perpendicular to the other two. Also P consists of two dumb or two bell shaped in contact with each other and each dumb can contain an electron and these two dumb are meeting head to head at a point where the electron is difficult to exist (zero electron density).



4- Sublevel (d) has 5 orbitals **(m = 2 x 2+1 = 5)**

Therefore, the level (d) has five directions

5- Sublevel (f) has 7 orbitals **(m = 2 x 3+1 = 7)**

Therefore, the level (f) has seven directions

4-Spain Q number (m)

Any orbital contain two electrons each electron spain around its axis during orbits around nucleus . Althought the electrons in the same orbitals carry the same negative charge we might exept them to repel. Yet due to the spain of electron around its axis a magnatic field will be arised so one electron spins around its axis clockwise while the other electron spins anti clockwise in order to from 2 opposite magnetic fields to decrease the force of repulsion between them which keep the atom stable.

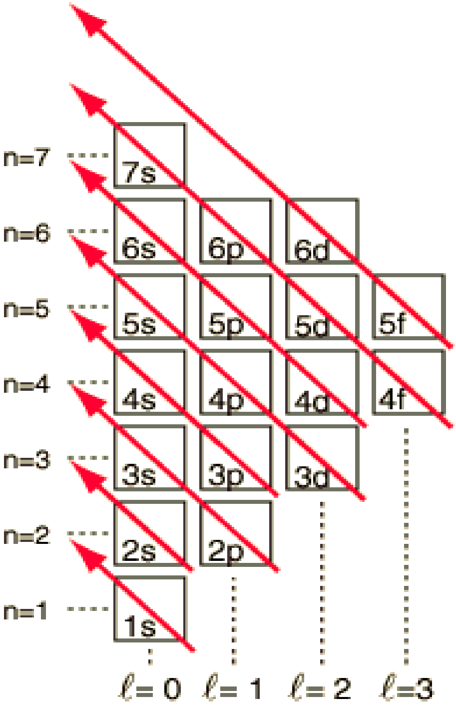
Spain Q number **is used to d**etects the direction in which the electron spins around its axis during its rotation around the nucleus

Principles of distributing electrons

There are two important rules which must be considered in distributing electrons in the atom. These rules are

1-Building-up principle

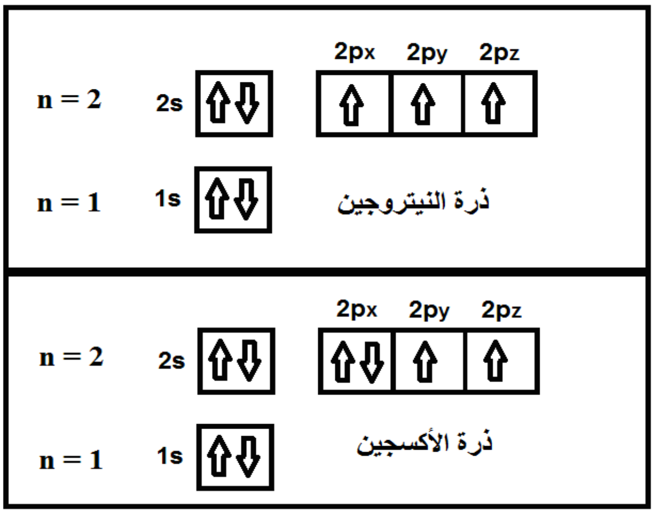
It states that electrons must fill the lower energy sub-levels with lower (n+L) first and then the higher-energy sub-levels.and if we have two orbital with same (n+L), the electron prefers to fill the orbital with lower(n)



1s ‹ 2s‹ 2p ‹ 3s ‹ 3p ‹ 4s ‹ 3d ‹ 4p ‹ 5s ‹ 4d ‹ 5p ‹ 6s ‹ 4f ‹ 5d ‹ 6p ‹ 7s ‹ 5f ‹ 6d ‹ 7p

2-Hund's Rule

No electron pairing takes place in a given sublevel until each orbital contains one electron



Chemical Bonding

Definition

- bond: forces that hold one atom to another in a compound
- To break a bond requires energy to be put in to overcome the forces of attraction
- Bond breaking is endothermic

- To make bonds causes a release of energy.
- *Bond making is exothermic*.
- Compounds have less energy (more stable) than the substances from which they form.
- Ex: water has less energy than the hydrogen and oxygen from which it formed.
- The energy stored in a bond is potential

- Compounds have less energy (more stable) than the substances from which they form.
- Ex: water has less energy than the hydrogen and oxygen from which it formed.
- The energy stored in a bond is <u>potential.</u>

Three Types of Bonds

1. Metallic
2. Ionic
3. Covalent
4. Hydrogen

Ionic Bonds

- Formed when <u>electrons are transferred between atoms.</u>

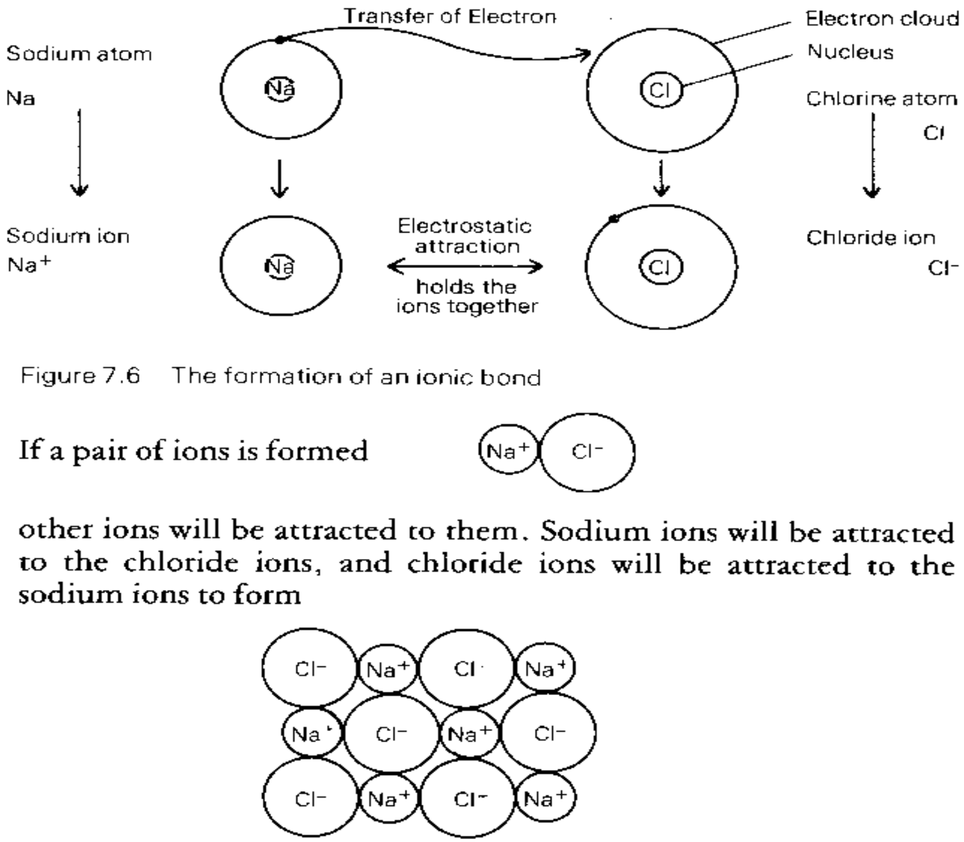


Figure 7.6 The formation of an ionic bond

If a pair of ions is formed

other ions will be attracted to them. Sodium ions will be attracted to the chloride ions, and chloride ions will be attracted to the sodium ions to form

Ionic Bonds: Metals and Non-metals

- Metals… lose e-.
- Non-metals… gain e-.
- They can exchange electrons to form a bond.

The Octet Rule

- Atoms are happy with a full valence shell.
- This achieved by gaining or losing electrons.

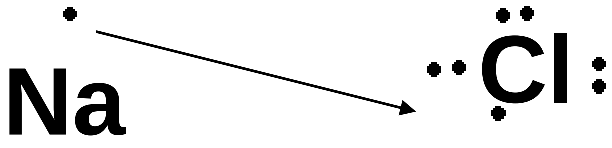
- Metals tend to have low ionization energies, so they lose e- easily.

- Non-metals tend to have high electronegativities, so they gain e’ readily.

To add to notes…

- The two Elements involved in an ionic bond have a **difference in electronegativity** (E.N.D.) that is greater than or equally to **1.7**
- Ex: NaCl E.N.D.= 2.3 (How do you find this? Look up electronegativity of each element on Table S and subtract.)
- Na = 0.9 Cl = 3.2
- 3.2 – 0.9 = **2.3**

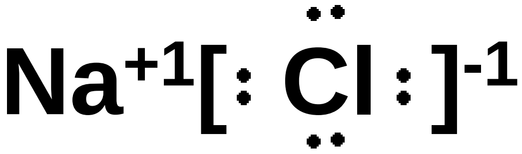
Example:
sodium reacts with chlorine



Has one valence electron, wants to get rid of it so that its valence shell is full

Has 7 valence electrons, needs one more to have a full valence shell

This reactions makes sodium chloride or table salt.

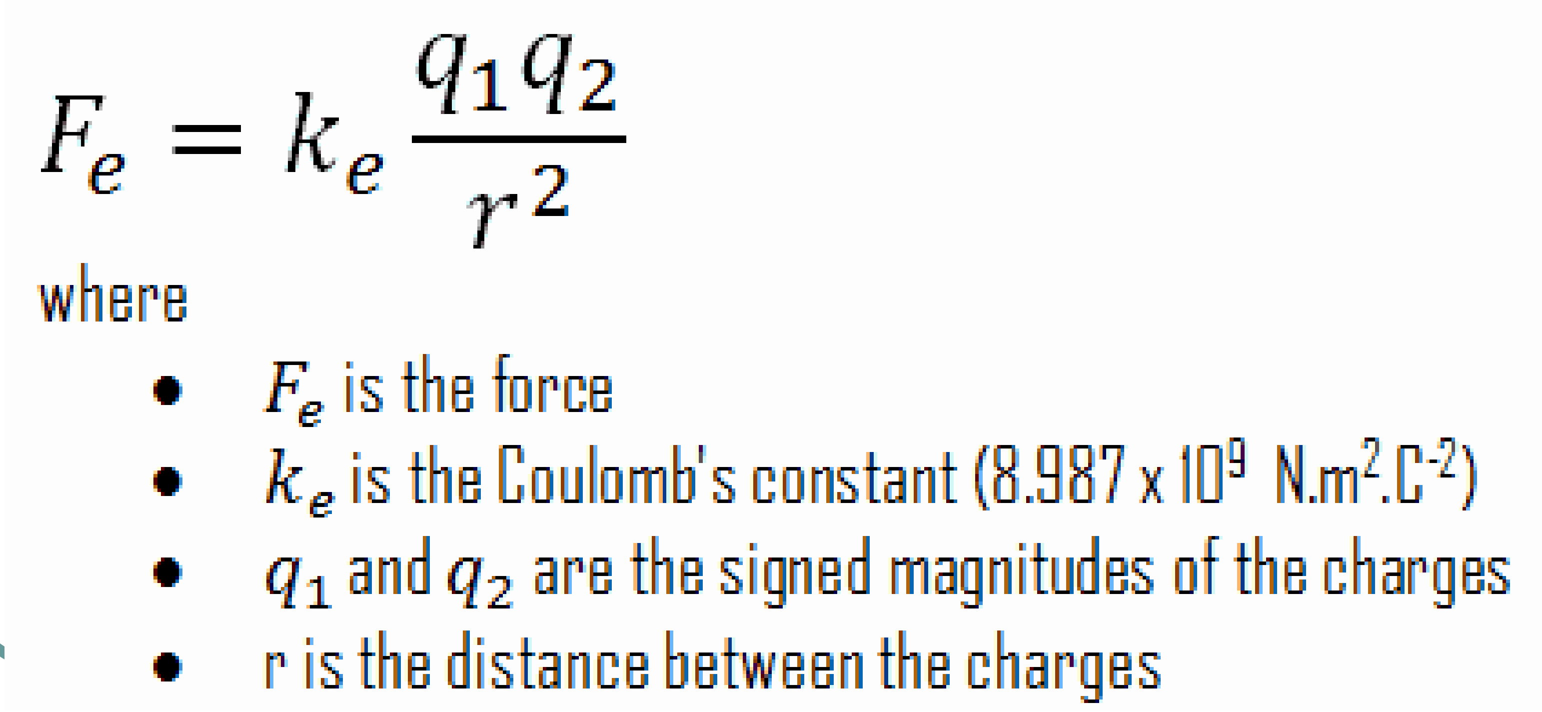


$Na^{+1}[:\ddot{\underset{..}{Cl}}:]^{-1}$

Summary of Concepts: Ionic Bonds

- Occur between **metals (+) and non-metals (-)**
- Involve a transfer of electrons from the metal to the nonmetal.
- The two elements have an **electronegative difference** of **1.7 or greater**
- We call ionic compounds **“salts”**

Coulomb's law can be used to calculate the energy of an interaction between a pair of ions. It can be used for both attractive and repulsive forces.



$$F_e = k_e \frac{q_1 q_2}{r^2}$$

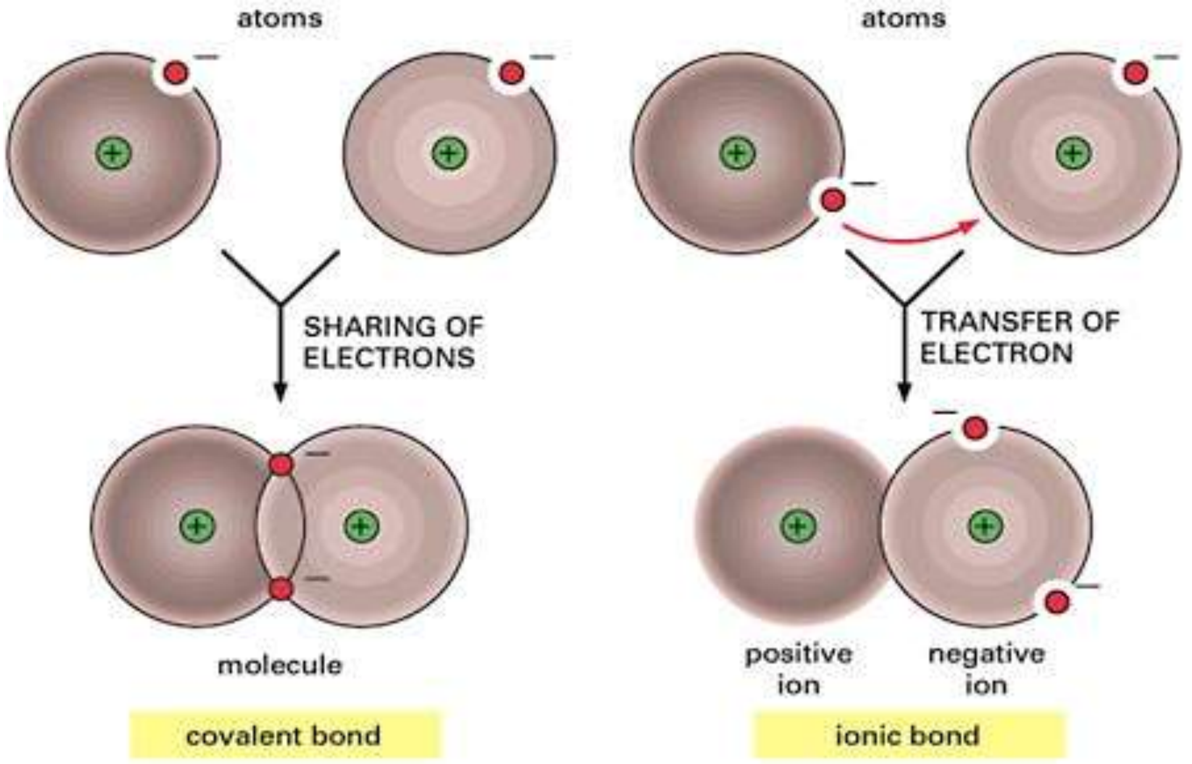
where

- F_e is the force
- k_e is the Coulomb's constant (8.987 x 10^9 N.m^2.C^{-2})
- q_1 and q_2 are the signed magnitudes of the charges
- r is the distance between the charges

Summary of Coulomb's Law

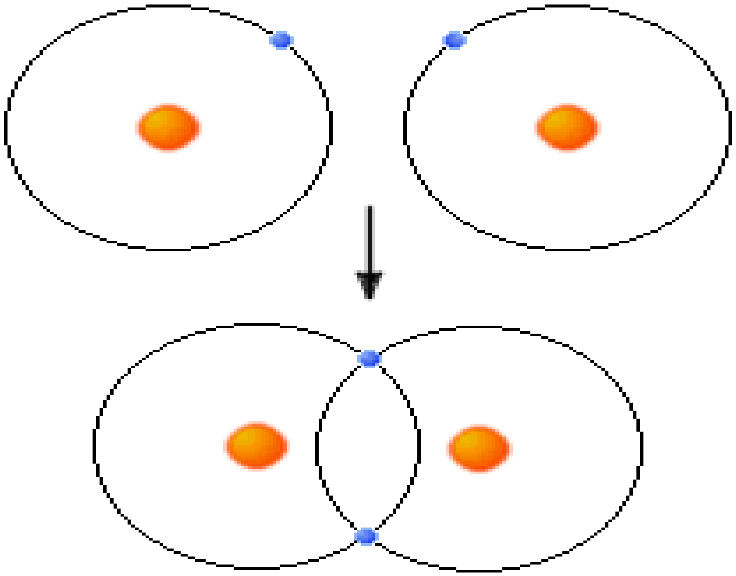
- The **strength of an ionic bond** can be determined by 2 factors:
  - **The charge of the ions**, the higher the numerical charge value the stronger the attraction.
  - **The distance of the ions**, the closer they are the stronger the bond.

Covalent Bonds vs. Ionic Bonds



Two Hydrogen Atoms

- The valence shells overlap and the electrons are shared making a more stable molecule.

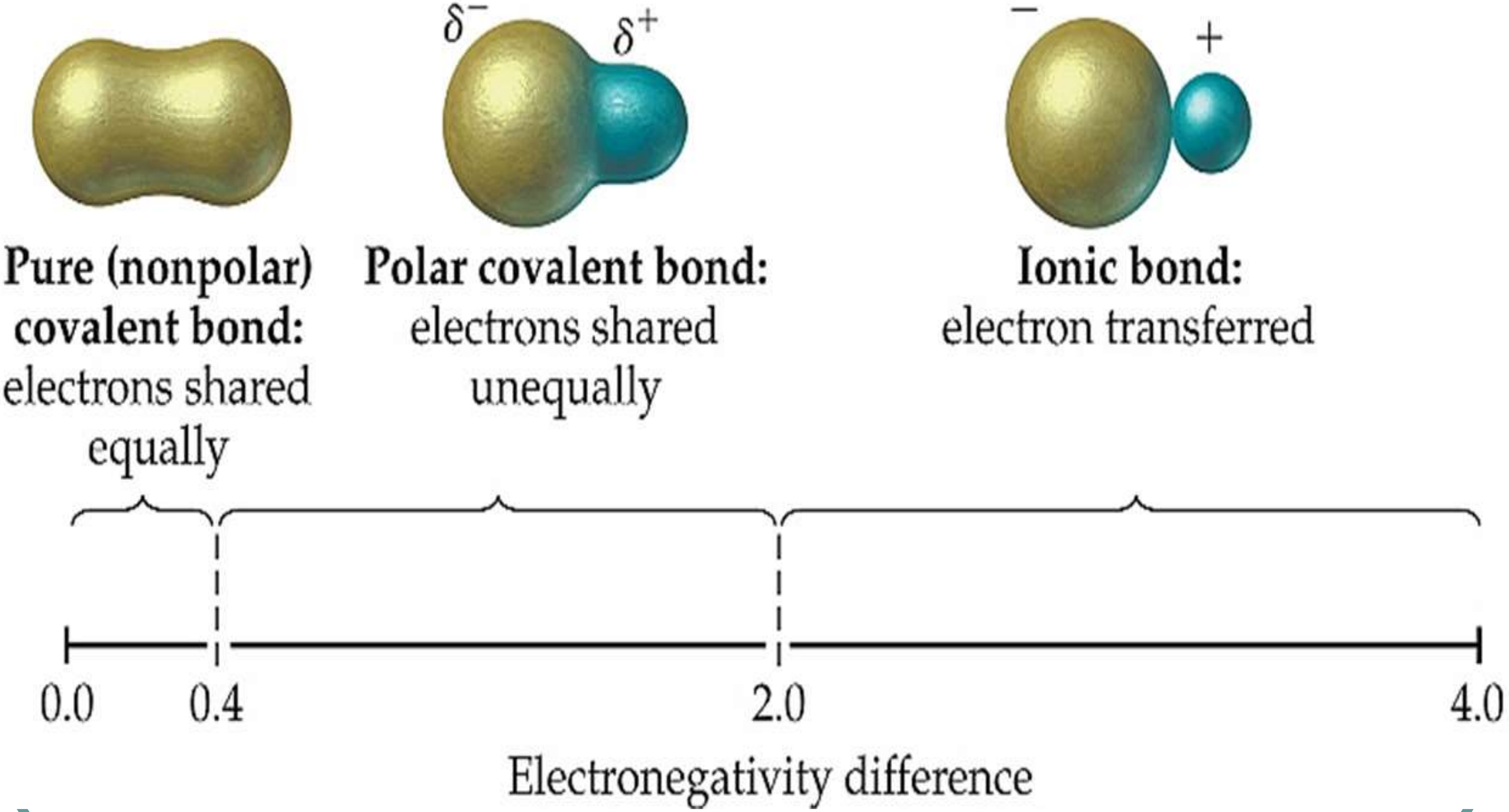


Covalent bond

Covalent Bonds

- Involves **SHARING** electrons because both elements have high **electronegativities**.
- Sharing of electrons can be equal (**non-polar**) or unequal (**polar**)
- Usually is between two **NON-METALS** (ex. H, C, O, N…)
- Covalent bonding like Ionic bonding results in a more stable compound, because the atoms involved meet the “octet rule”.

Types of Covalent Bonds



Pure (nonpolar) covalent bond: electrons shared equally

Polar covalent bond: electrons shared unequally

Ionic bond: electron transferred

Bond Polarity

Polar Covalent Bonds (between two atoms)

- Covalent bonds unlike ionic bonds, are NOT composed of oppositely charged ions.
- However, we find that since the electrons are **shared** and each element has a different **electronegativity**, the shared electrons are often **shared unequally**.
- Meaning the electrons spend more time around the more electronegative atom.
- This causes one atom to be slightly negative and one atom to be slightly positive.
- This is known as a **dipole** moment. One positive end and one negative end 2 oppositely charged poles.

Bond Polarity

Non-polar Covalent Bond (between two atoms)

- If both of the atoms involved in a bond have the same electronegativity then the electrons are shared **equally**.
- This will happen if a bond forms between two of the same atoms. Since the atoms are the same they will have the same electronegativity and share the electron(s) **equally**.
- When electrons are shared equally their will never be a dipole moment and the bond will be **non-polar**.

Molecule Polarity

Polar Molecule

- A molecule that is **polar** (also known as a **dipole**) is asymmetrical which results in an uneven distribution of charge throughout the entire molecule.
- Ex:

H_2O which has a bent shape and is assymetrical

Molecule Polarity

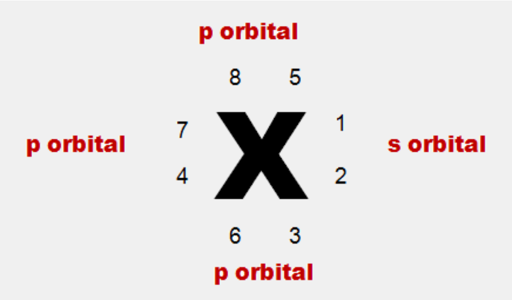
Non-Polar Molecules

- A molecule that is non-polar is symmetrical in shape which results in an even distribution of charge.
- Ex

CH_4 which is tetrahedral symmetrical in shape

Octet Rule

- Most elements follow the octet rule which states that atoms tend to gain, lose, or share electrons until they are surrounded by eight valence electrons.
- An octet of electrons consists of full s and p subshells in an atom.
- Elements that do not follow the octet rule include hydrogen and helium. They follow the duet rule.

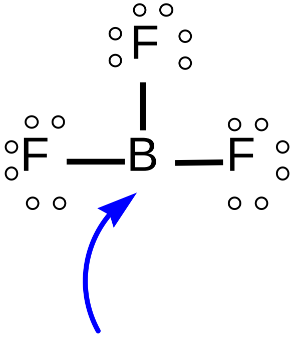


Drawing Lewis Dot Diagrams for octet rule

1) Add up all the valence electrons for an atom in the molecule.
2) The “central atom” is the one that is most electronegative or there will only be one of this atom.
3) Add the electrons until you reach the total remembering that all the atoms must obey the octet rule except for hydrogen which obeys the duet rule.
4) Also remember that one pair of electrons between two atoms is a single bond two is a double and three pairs is a triple bond. These are known as **bonded or shared pairs** of electrons.
5) Not every electron has to be bonded. We call these the **lone or non-bonded pairs** because they are “lonely”.

4.3 Exceptions to the Octet Rule

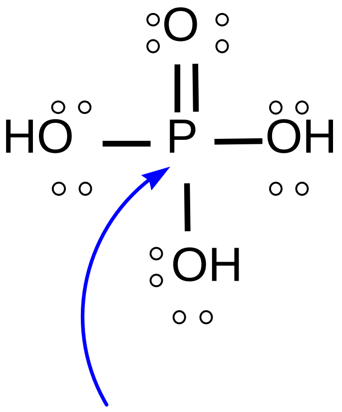
- **Most of the common elements generally follow the octet rule.**
- **H is a notable exception, because it needs only 2 e^- in bonding.**
- **Elements in group 3A do not have enough valence e^- to form an octet in a neutral molecule**.



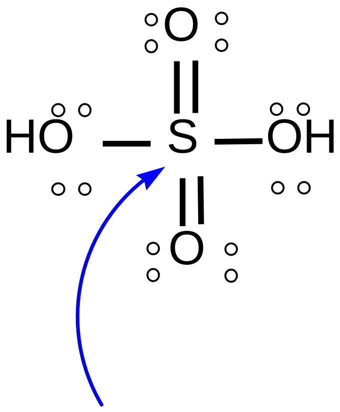
only 6 e^- on B

4.3 Exceptions to the Octet Rule

- **Elements in the third row have empty *d* orbitals available to accept electrons.**
- **Thus, elements such as P and S may have more than 8 e⁻ around them.**



10 e⁻ on P



12 e⁻ on S

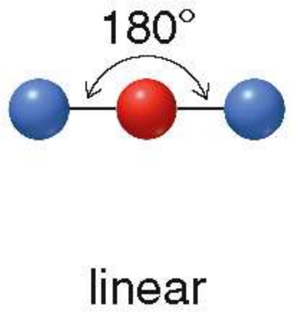
4.6 Molecular Shape

- **The Lewis structure gives information about how the atoms are connected, but it implies nothing of the geometry or shape.**
- **To determine the shape around a given atom, first determine how many groups surround the atom.**
- **A group is either an atom or a lone pair of electrons.**
- **Use the VSEPR theory to determine the shape.**
- **The most stable arrangement keeps the groups as far away from each other as possible.**

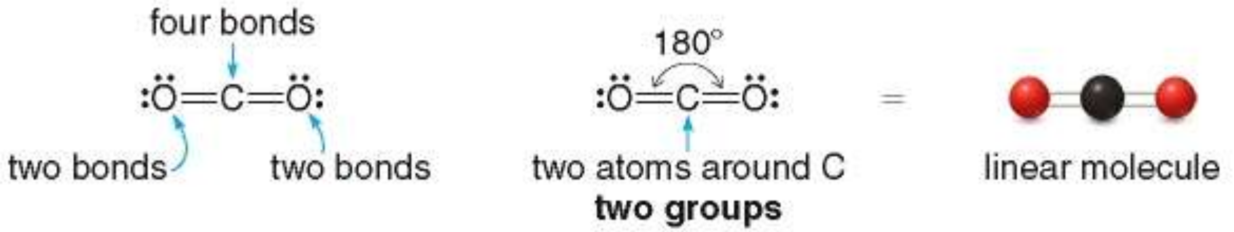
4.6 Molecular Shape

A. Two Groups Around an Atom

- **Any atom surrounded by only two groups is linear and has a bond angle of 180º.**



- An example is CO_2:

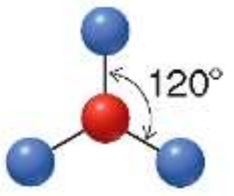


- **Ignore multiple bonds in predicting geometry. Count only atoms and lone pairs.**

4.6 Molecular Shape

B. Three Groups Around an Atom

- Any atom surrounded by three groups is trigonal planar and has bond angles of 120º.



trigonal planar

- An example is H_2CO (formaldehyde):

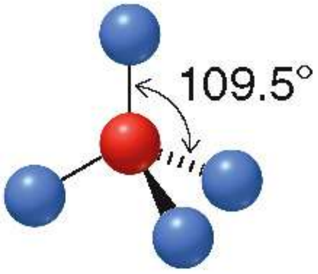
three atoms around C
three groups

trigonal planar
molecule

4.6 Molecular Shape

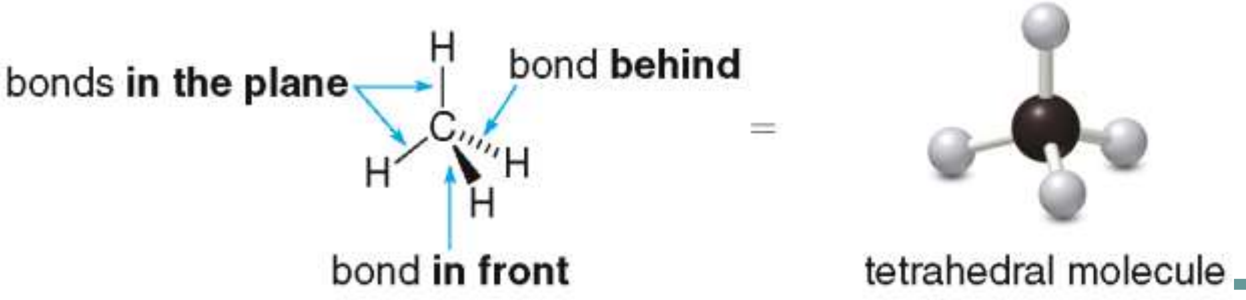
C. Four Groups Around an Atom

- Any atom surrounded by four groups is tetrahedral and has bond angles of 109.



tetrahedral

- An example is CH_4 (methane):

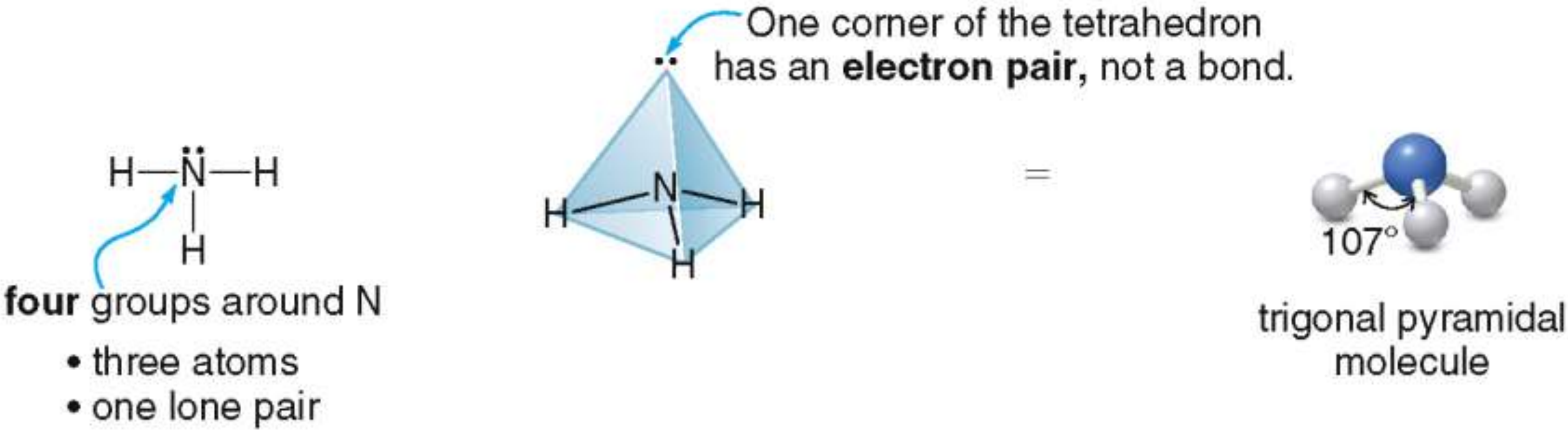


tetrahedral molecule

4.6 Molecular Shape

C. Four Groups Around an Atom

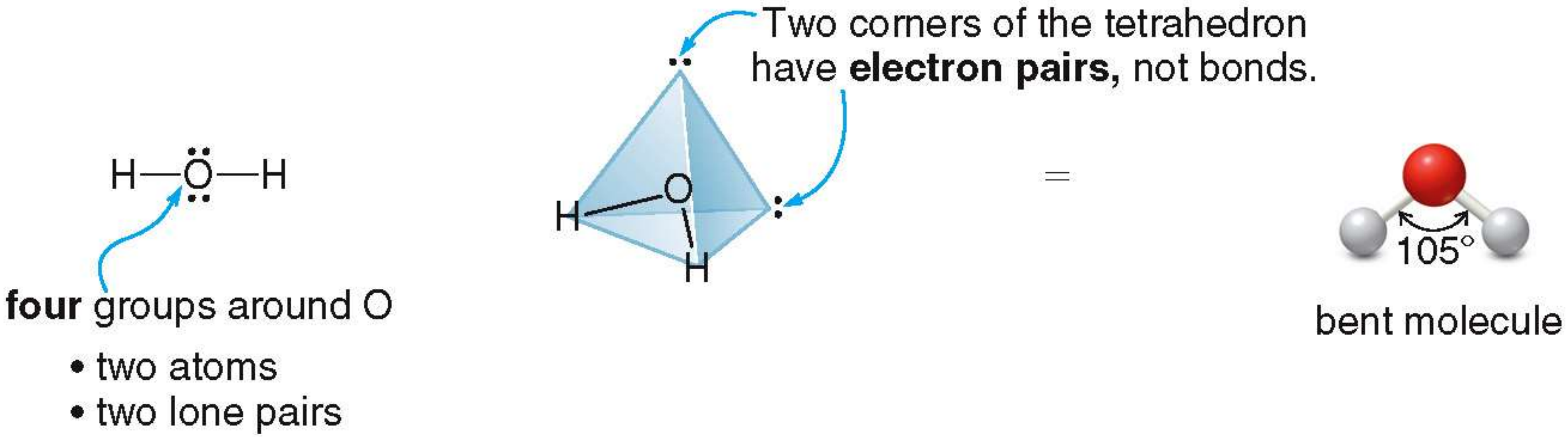
- If the four groups around the atom include one lone pair, the geometry is a trigonal pyramid with bond angles of 107^{o}, close to 109.5^{o}.
- An example is NH_3 (ammonia):



4.6 Molecular Shape

C. Four Groups Around an Atom

- If the four groups around the atom include two lone pairs, the geometry is bent and the bond angle is 105° (i.e., close to 109.5°).
- An example is H_2O:



4.6 Molecular Shape

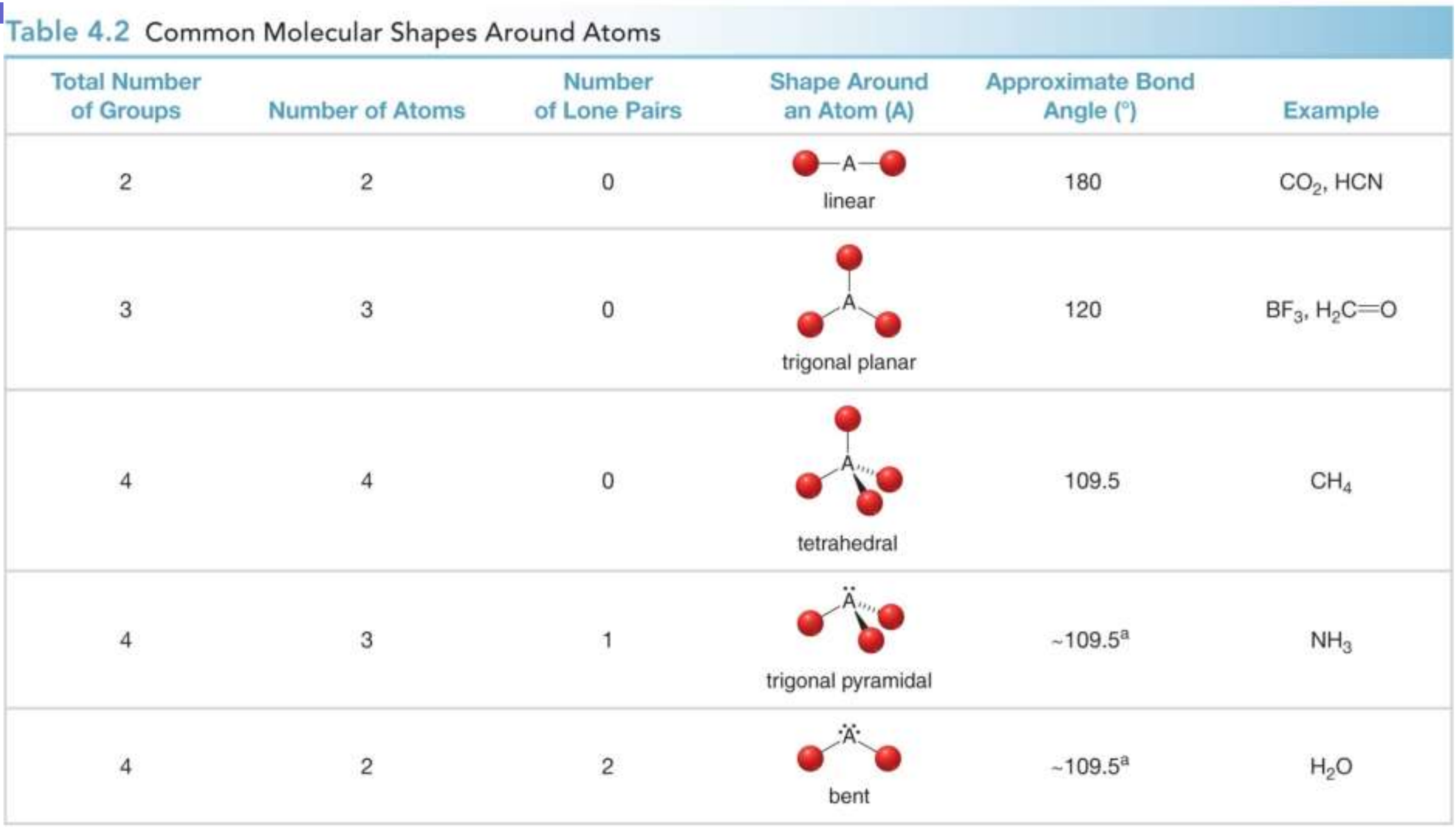


Table 4.2 Common Molecular Shapes Around Atoms

Total Number of Groups	Number of Atoms	Number of Lone Pairs	Shape Around an Atom (A)	Approximate Bond Angle (°)	Example
2	2	0	linear	180	CO_2, HCN
3	3	0	trigonal planar	120	BF_3, $H_2C{=}O$
4	4	0	tetrahedral	109.5	CH_4
4	3	1	trigonal pyramidal	~109.5[a]	NH_3
4	2	2	bent	~109.5[a]	H_2O

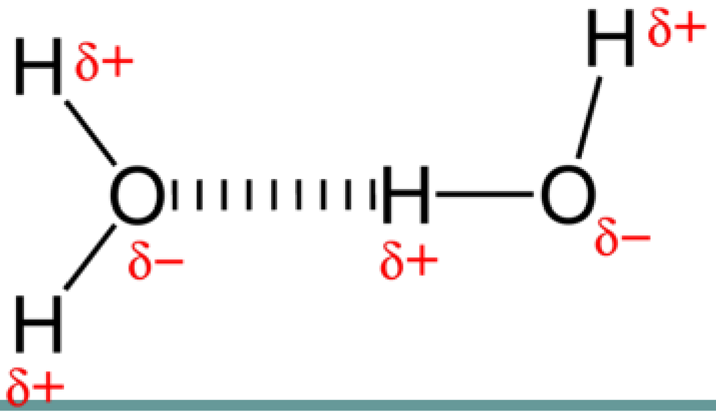
[a] The symbol "~" means approximately.

Polar Molecules

- **Hydrogen bonding** is one type of intermolecular force it is formed between a hydrogen atom in one molecule and a nitrogen, oxygen, or fluorine atom in another molecule.
- Hydrogen bonding is the **strongest** intermolecular force for all molecular substances.
- Substances with H-bonding have relatively **high boiling points**, because the force is so strong.
- The most important example of this is found in water. The strength of Hydrogen bonding is the reason why water has a relatively high boiling point compared to other molecular compounds.

Polar Molecules (Continued)

- If water did not have H-bonding it would not exist on earth as a liquid only a vapor. The diagram below shows how the H-bonding occurs.
- H-bonding also occurs between **ammonia molecules and Hydrogen Fluoride molecules**.



Polar Molecules (Continued)

- All other intermolecular forces for polar molecules are due to the **dipole-dipole attractions** between the molecules, and the force is not quite as strong as Hydrogen bonding.
- Remember H-Bonding only occurs between the **HYDROGEN** atom of one molecule and either an **OXYGEN, NITROGEN or FLUORINE** atom of another molecule.

Properties of Bonds

- Metallic, covalent and ionic bonds all have varying **properties**.
- These properties can be used to identify the bonding of **unknown substances**.
- The three main properties that can be used to distinguish bond type are **Melting & Boiling Points**, **Hardness** in the solid state, and **Conductivity** in the solid, liquid and aqueous states.

Properties of Metallic, Ionic and Covalent Bonds Summary

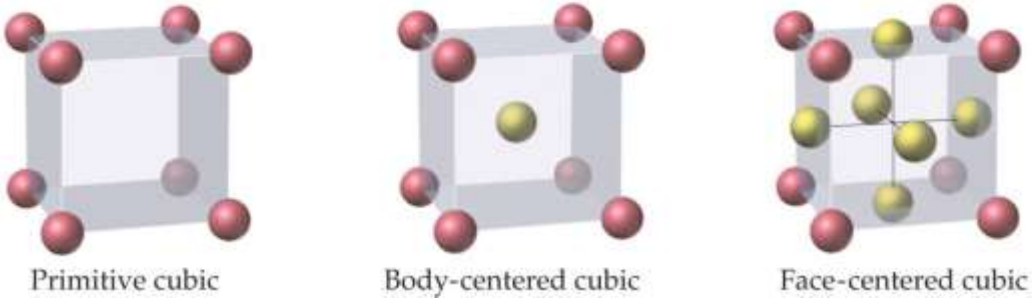
Bond Type	Melting & Boiling Points	Hardness	Conductivity		
			Solid	Liquid	Aqueous
Metallic	High (except Hg)	Hard	Yes	Yes	Yes
Covalent	Low	Soft	No	No	No
Ionic	High	Hard	No	Yes	Yes

Properties of Metallic, Ionic and Covalent Bonds Summary

Bond Type	MP & BP	Hardness	Conductivity		
			Solid State	Liquid State	Aqueous
Metallic	High (except Hg)	Hard	Yes	Yes	Yes
Ionic	High	Hard	No	Yes	Yes
Covalent	Low	Soft	No	No	No

Ionic Solids

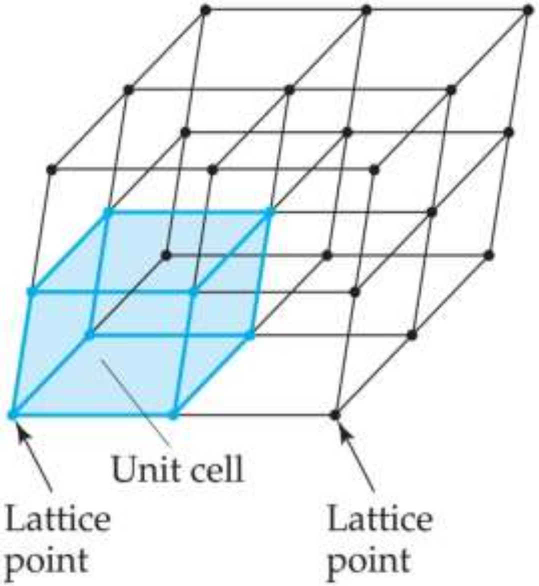
- Repeating particles in crystals are ions (+/-).
- **Examples:** NaCl, KNO_3, $CaCl_2$, etc.
- Diagram



Primitive cubic

Body-centered cubic

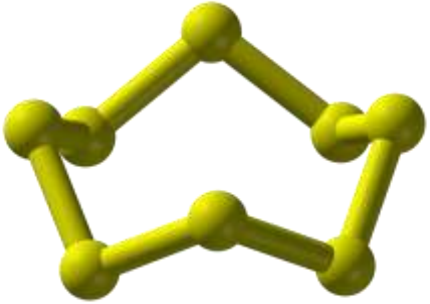
Face-centered cubic

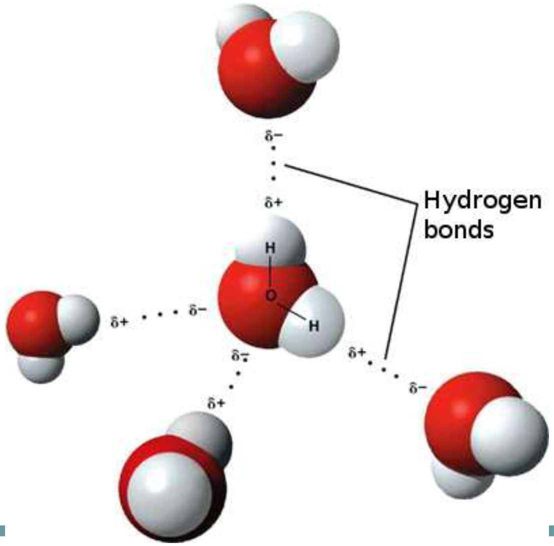


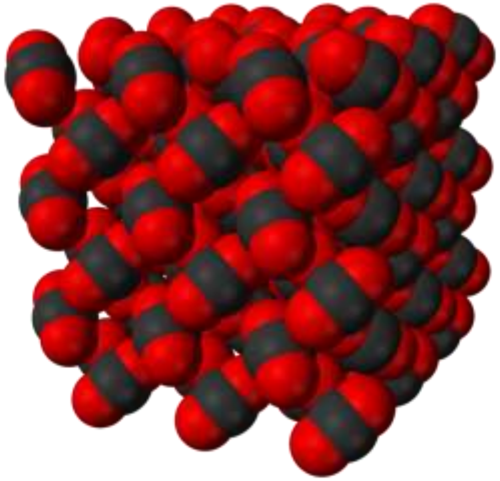
Ionic Solids

- Properties:
  - High melting point (strong bonds)
  - Electrically conductive ONLY when dissolved in water or in liquid form (mobile charged particles)
  - Water soluble (exceptions on Table F)
  - Relatively hard

- **Molecular Solids:** Repeating particles are atoms or molecules.
- Examples: CO_2, H_2O, S_8
- Diagram:







Molecular Solids

- Properties:
  - Relatively low MP and BP (based on Intermolecular forces)
  - Soft
  - Non-conducting
  - Only polar molecules are soluble in water

Covalent Network Solids

- Properties:
  - Very strong covalent bonds
  - Very high melting point
  - Very hard
  - Not soluble in water

Ionic Bonding:
What does the word "ionic" mean?

- Comes from the word ion meaning charged particle.

- Ions are formed when atoms gain or lose e-.

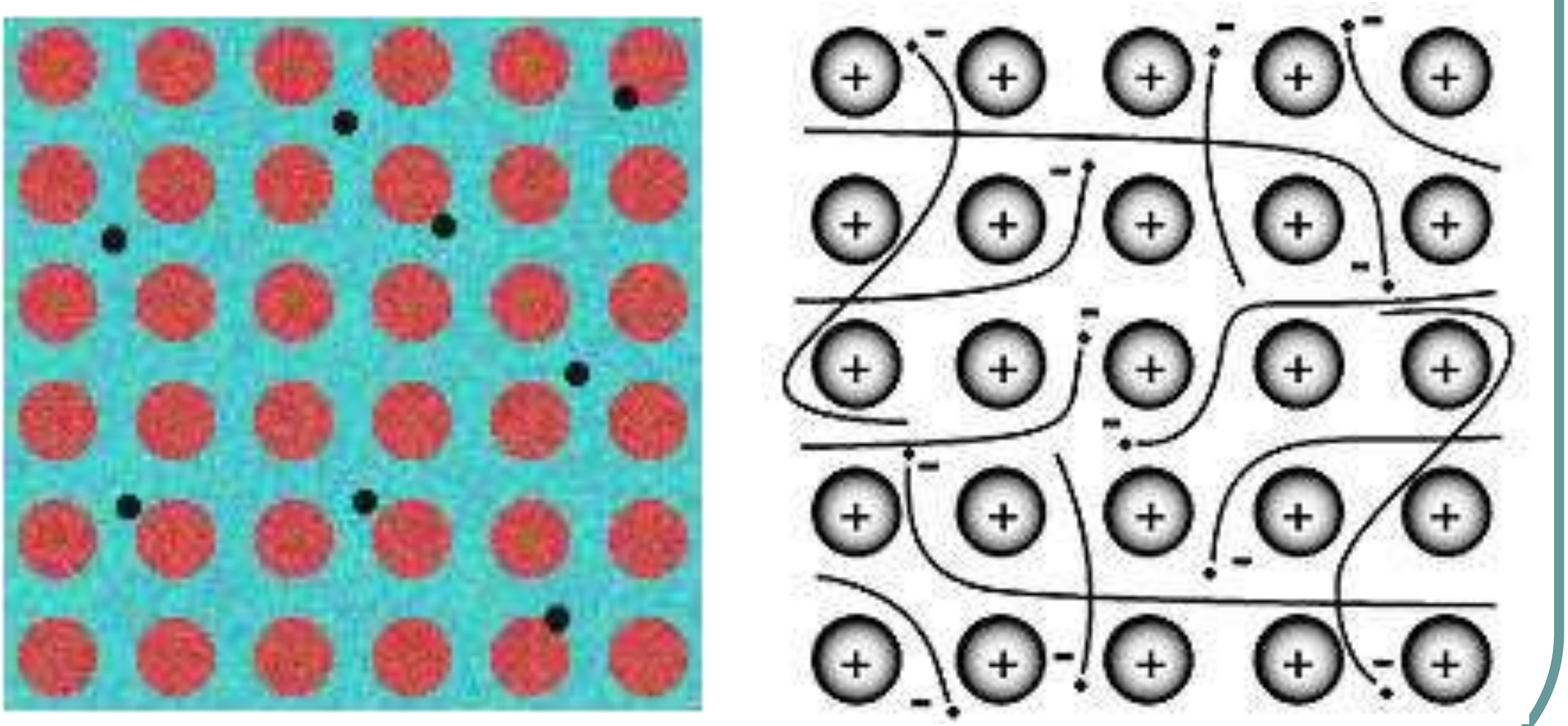
Metallic Bonds

- **Definition:** bonds between atom in a metal; ions held together in a crystalline lattice in a "sea of mobile electrons"

Metallic Bonds

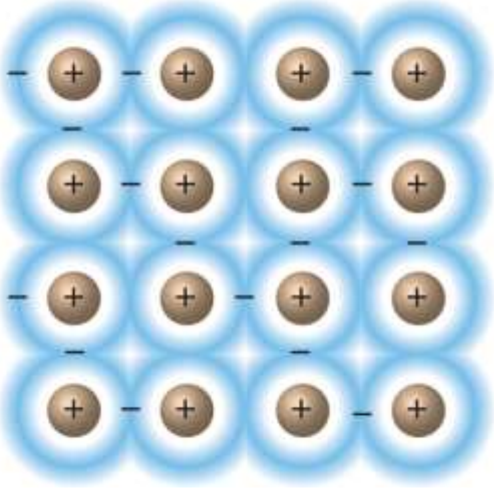
- Conduct electricity because of <u>freely moving electrons</u>.
- Any substance that has moving charged particles (typically either mobile electrons or ions) will conduct electricity.
- High <u>MP</u> and high <u>BP</u> because <u>bonds are strong.</u>
- malleable

Diagrams



Metallic Solids

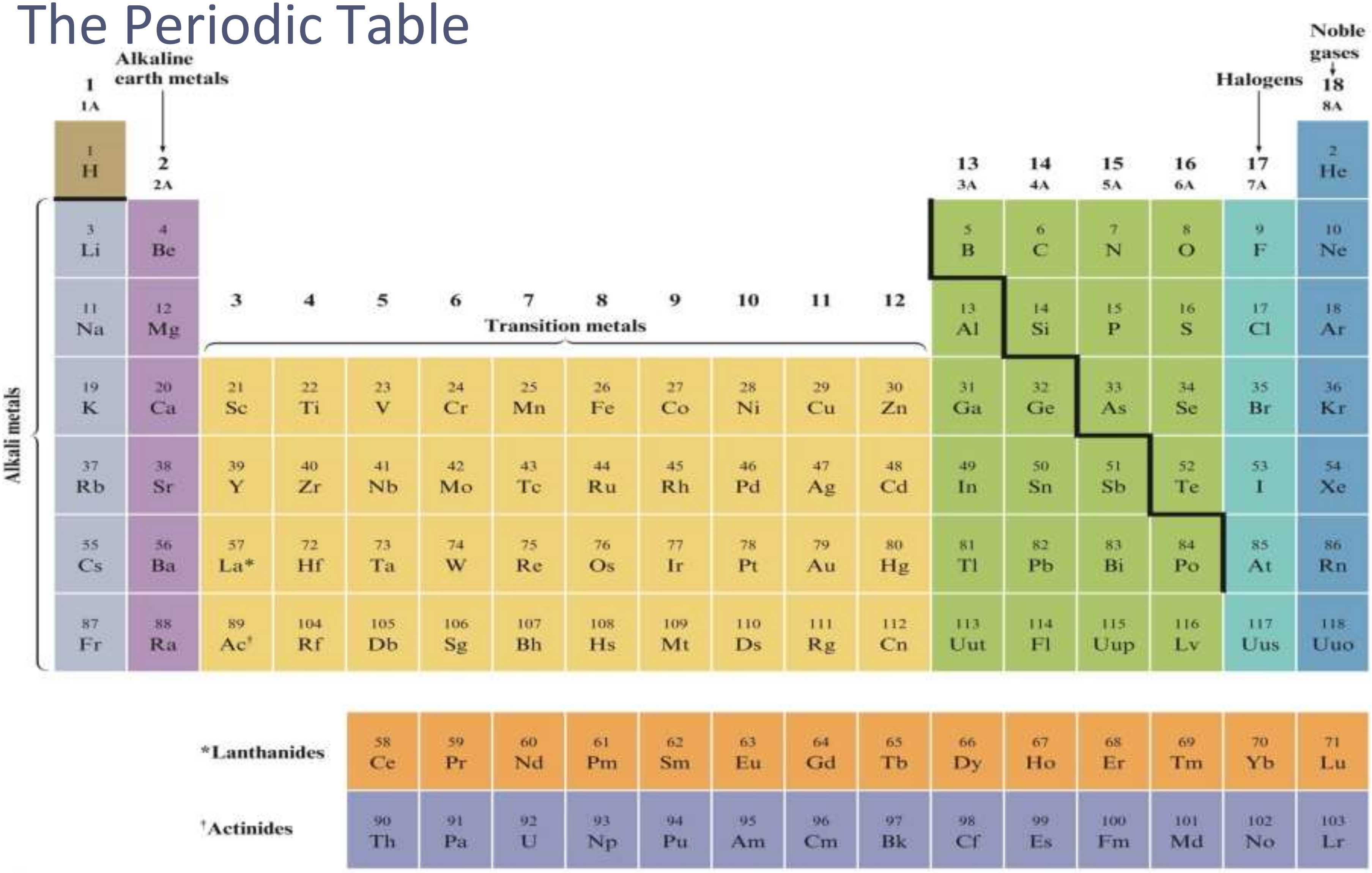
- Repeating particles are <u>metal atoms</u> with mobile valence electrons moving throughout the crystal; a “sea of mobile electrons”
- Examples: Ag, Au, Na, Cu, Zn
- Diagram



Metallic Solids

- Properties:
  - Some are hard, others are soft
  - Conduct electricity and heat
  - Malleable/ductile
  - insoluble

The Periodic Table



Alkali metals

Alkaline earth metals

Transition metals

Halogens

Noble gases

1 1A	2 2A	3	4	5	6	7	8	9	10	11	12	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

*Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
†Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

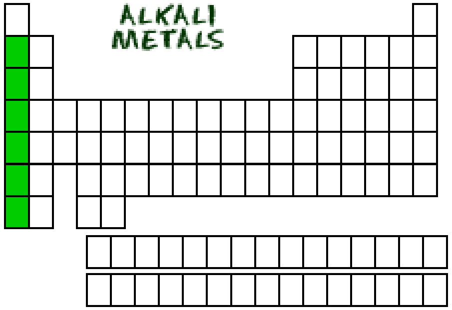
The Periodic Table

- *Metals vs. Nonmetals*
  - Metals
    - Majority of elements
    - Good conductors of heat/electricity
    - Solids at room temperature (except Mercury)
    - Ductile, malleable
    - Vary in reactivity
    - Transition Metals (Gr 3-12)
      - Bridge between sides of table

- Nonmetals
  - Opposite to metals
  - Poor conductors, brittle
  - Gases at room temp (except bromine, iodine, sulfur, selenium, phosphorus, and carbon)
  - Vary in reactivity
- Metalloids
  - Properties of metals and nonmetals
  - Variable conductivity
  - B, Si, Ge, As, Sb, Te, At

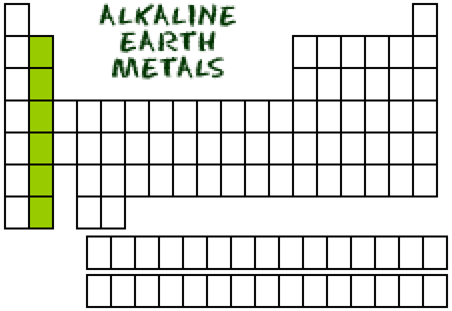
- *Periods* – horizontal rows of elements
    - 7 periods on table
    - Increasing atomic number
- *Groups or Families*
    - elements in the same vertical columns
    - have similar chemical properties
    - arranged by the number of electrons in the outermost energy level (valence electrons)

Alkali Metals



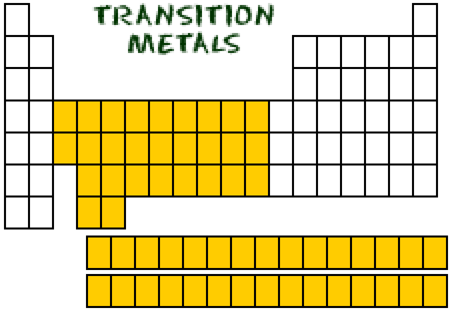
- Group 1A
- Metals with a single valence electron
- Very reactive- only found in nature as compounds
- Extremely reactive with water
- Reactivity increases from top to bottom
- Francium is most reactive metal
- Most are stored in oil or in argon gas to protect them from violent reactions

Alkaline Earth Metals



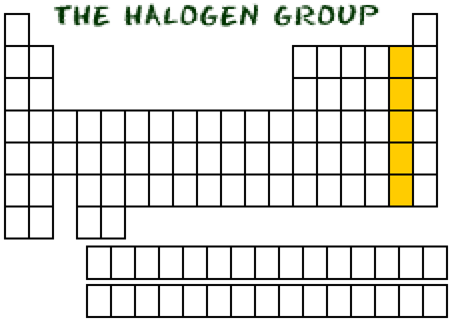
- Group 2A
- 2 valence electrons
- Harder than metals in Group 1
- Variable reactivity with water
- Magnesium- assists in photosynthesis, lightweight metal mixtures
- Calcium- bones, teeth, chalk, limestone, coral, pearls

Transition Metals

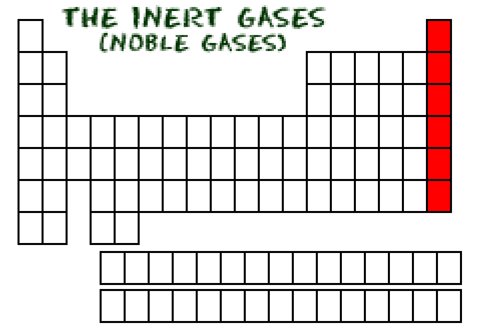


- Groups 1B-10B (or 3-12)
- All metals
- Bridge between 2A and 3A (or 2 and 13)
- Varying number of valence electrons
- Varying reactivity
- Gold, Silver, Copper, and other valuable metals here

The Halogens



- Group 7A
- 7 valence electrons
- Fluorine, Chlorine, Bromine, Iodine- nonmetals
  - Fluorine is most active nonmetal
    - toothpaste, Teflon
  - Chlorine is used in bleach and to kill bacteria in pools and drinking water
  - Iodine keeps the thyroid gland functioning (metabolism)
- Astatine- metalloid



The Noble Gases

- Group 8A
- 8 valence electrons (except He, which has 2)
- Outermost energy level considered “full”, so they do not react
- Used to help other things remain unreactive, such as silicon in computer chips and alkali metals
- Used in light bulbs to help filaments last longer, and in “neon” lights
- Helium, Neon, Argon, Krypton, Xenon, Radon- all nonmetals

The Periodic Table