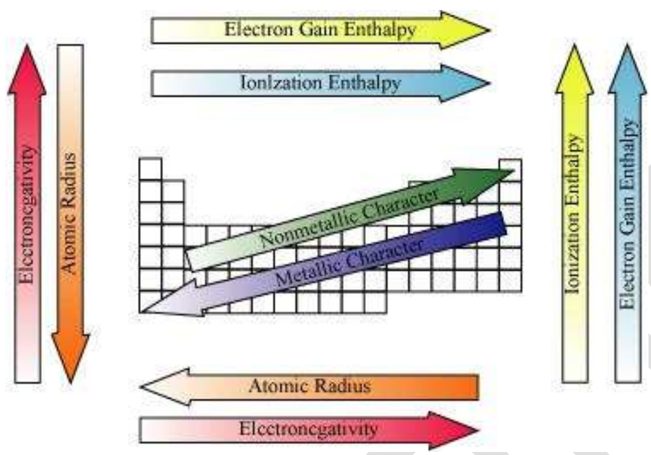
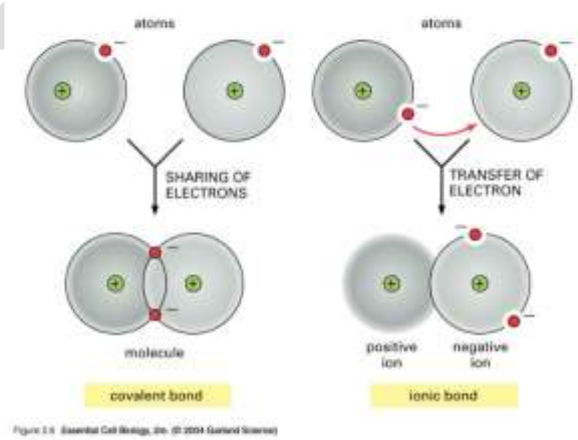
Unit 6:

The Periodic Table & Bonding





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Class Period: ________

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Unit 6 Vocabulary:

1. Alkali metal: An element in Group 1 of the periodic table that is extremely reactive.
2. Alkaline earth metal: An element in Group 2 of the periodic table that is very reactive.
3. Anion: A negatively charged ion.
4. Atomic radius: The size of an atom. Sometimes called the "covalent atomic radius".
5. Brittle: The ability to be crushed into pieces if hammered; a property of nonmetals.
6. Cation: A positively charged ion.
7. Diatomic molecule: A nonmetal that forms one (or more) nonpolar covalent bonds with another atom of the same element to form a molecule. Diatomic molecules form when there are no other types of elements to readily bond with. Diatomic molecules include: Br_2, I_2, N_2, H_2, Cl_2, O_2, and F_2.
8. Ductile: The ability to be stretched into a wire; a property of metals.
9. Dull: The inability to reflect light; a property of nonmetals.
10. Group: Columns (vertical) on the period table with elements that have the same number of valence electrons and similar chemical properties.
11. Halogen: An element of Group 17 of the periodic table that is extremely reactive.
12. Ionic bond: A bond formed when a metal atom loses its valence electron(s) to a nonmetal atom, forming positive and negatively charged ions that are mutually attracted to each other.
13. Ionic radius: The size of an ion compared to the original atom. Metal atoms lose electrons forming + ions that are smaller than the original atom. Nonmetal atoms gain electrons forming - ions that are larger than the original atom.
14. Luster: The ability to reflect light; a property of metals.
15. Malleable: The ability to be hammered or rolled into thin sheets; a property of metals.
16. Metallic bond: A bond formed between metal atoms of the same element resulting from the atoms losing electrons to each other and sharing the electrons loosely.
17. Metalloid: An element that exhibits properties of both metals and nonmetals.

18. Molecular orbital: A hybrid orbital made up of the shared unpaired valence electrons of two nonmetallic atoms. This orbit belongs to both of the bonded atoms rather than to any specific atom.
19. Monoatomic molecule: An atom of a noble gas, which is considered to be a molecule as there are no unpaired electrons in a noble gas atom.
20. Noble gas: An element in Group 18 of the period table that is unreactive.
21. Nonmetal: Elements that have high electronegativity and ionization energy and a small atomic radius. Nonmetals tend to gain or share electrons when forming chemical bonds.
22. Nonpolar covalent bond: A bond formed between two nonmetal atoms when unpaired electrons of two atoms are shared equally. The electronegativity differences ranges between 0 and 0.4.
23. Nonreactive: Not capable of easily undergoing a chemical change.
24. Oxidation: The loss of valence electrons from an atom or an ion, resulting in the increase in oxidation number of an element.
25. Period: rows across (horizontally) the period table that denote elements with the same number of principal energy levels.
26. Polar covalent bond: A bond formed between two nonmetal atoms when unpaired electrons of two atoms are shared unequally. The electronegativity differences ranges between 0.5 and 1.7.
27. Reactive: Capable of undergoing a chemical change.
28. Reduction: The gain of valence electrons from an atom or an ion, resulting in the decrease in oxidation number of an element.
29. Semiconductor: An element that may act as either a conductor or an insulator, depending on the situation.
30. Stock System: A method of naming ions of elements that may form more than one possible positive charged ion. The Stock System uses Roman numerals after the ion name to denote the ions amount of positive charge.
31. Transition metal: An element in Groups 3-12 of the periodic table. Many transition metals have colored ions.

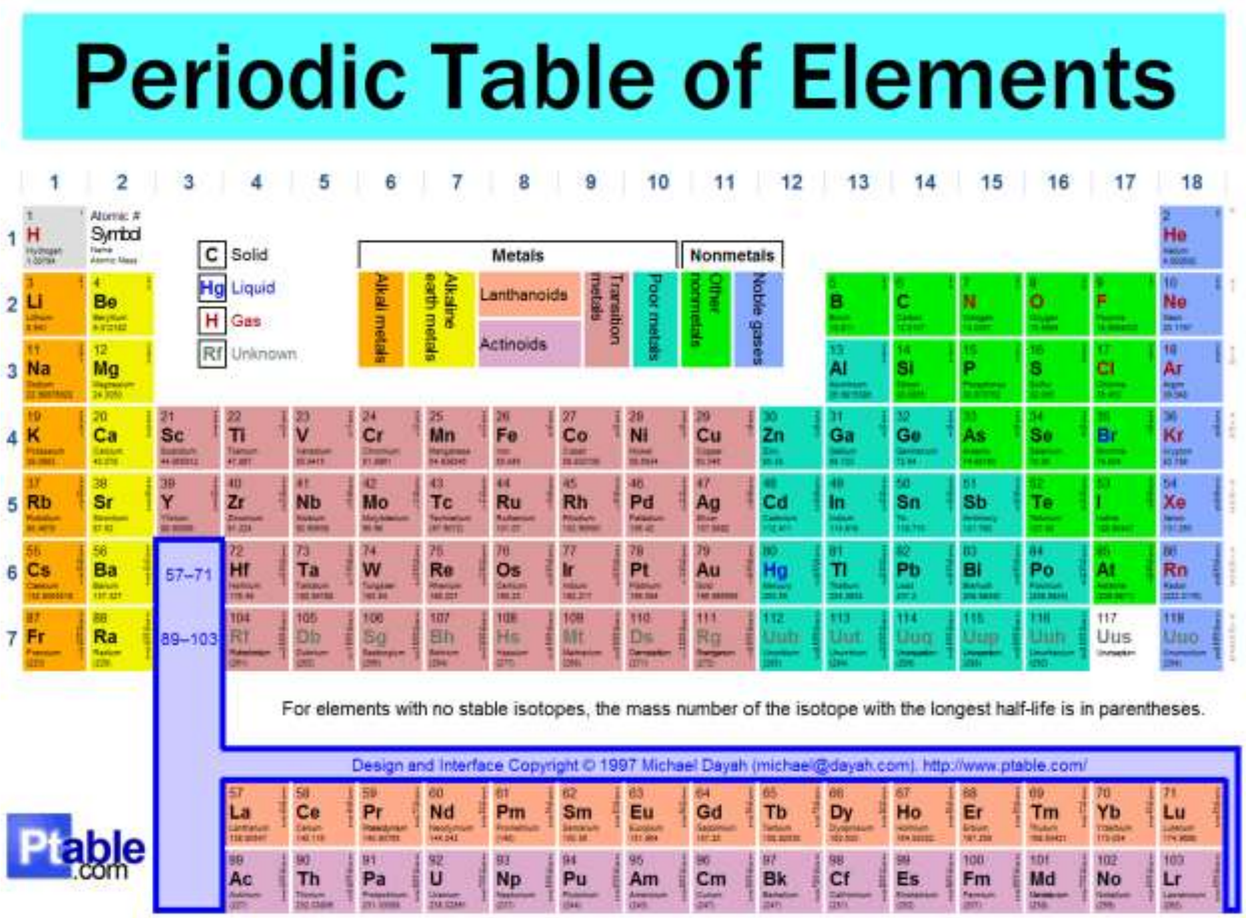
Notes page:

Unit 6 Homework Assignments:

Assignment:	Date:	Due:

Topic: **The Periodic Table**

Objective: How did the structure of the atom influence this design?



Development of the Periodic Table of Elements:

The modern Periodic Table of Elements was developed by Dmitri Mendeleev in 1869.

Modern Periodic Law:

Modern Periodic Law states that properties of elements are periodic functions of their atomic numbers. As atomic number increases sequentially (in order), certain properties such as valence electrons, ionization energy, and ion charge repeat periodically. Periodically means "at certain intervals" are repeated across horizontal rows. Elements are ordered in horizontal rows (periods) and vertical columns (groups).

Direction	Importance	Examples
PERIODS (rows)	All elements in the same period have the same number of principal energy levels in their atomic structure	Na, Mg, Al, Si, P, S, Cl & Al are all in Period 3. They all have three PELs in their atomic structure
Groups (columns)	All elements in the same group have the same number of valence electrons, therefore they lose or gain the same number of electrons, form similar chemical formulas and have similar chemical properties	Li, Na, K, Rb, Cs & Fr are all in Group 1. They all have one valence electron, they all lose the one valence electron when forming +1 ions, and they all are extremely reactive. Group 1 atoms have similar chemical properties and from the following formulas when bonding with oxygen: Li_2O, Na_2O, K_2O, Rb_2O, Cs_2O & Fr_2O

[Watch The Periodic Table: Crash Course Chemistry #4](https://www.youtube.com/watch?v=0RRVV4Diomg)

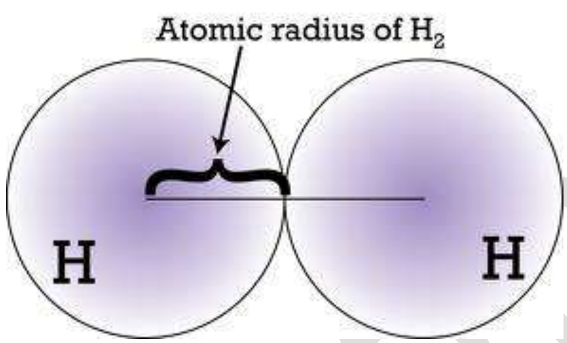
https://www.youtube.com/watch?v=0RRVV4Diomg

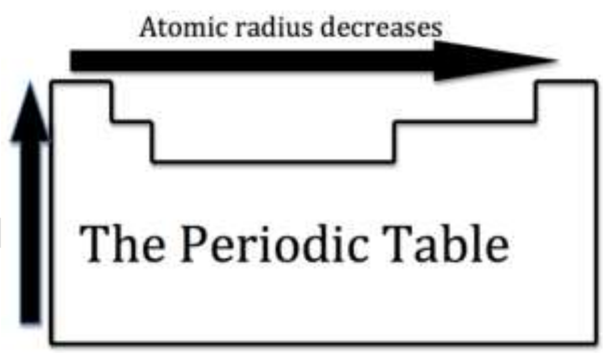
Topic: **Sizes of Atoms**

Objective: How does atomic radius change within the Periodic Table?

Sizes of Atoms:

Atomic radius: One-half (1/2) the measured distance between two nuclei of the same element while in the solid phase.

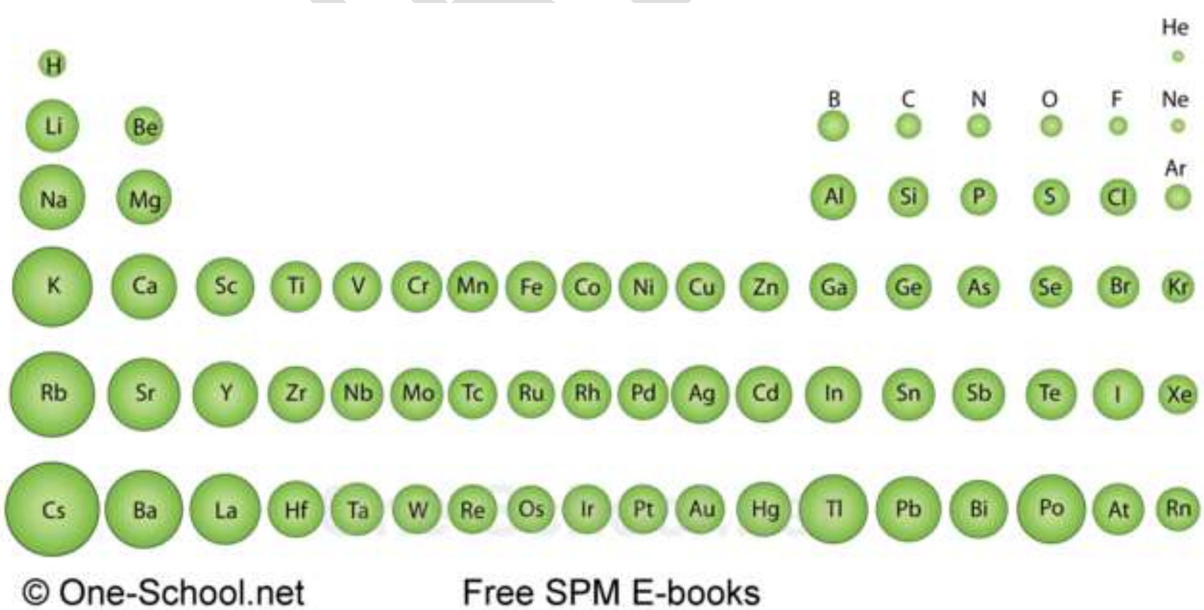




1. Within a **period** of the table, atomic **radius** generally **decreases** as the atomic number increases. This is due to an increase in nuclear proton (positive) charge as the atomic number increases. The increased **positive** nuclear charge **attracts** the negative valence **electrons** closer towards the nucleus, **decreasing** atomic **radius**.
   - As an example, for atoms in Period 3, all the valence electrons are in the same primary energy level (3 PELs). As the atomic number increases, the number of protons increases. Sodium (Na) has 11 protons, magnesium (Mg) has 12 protons, and aluminum (Al) has 13 protons, and so on. As the number of protons increases, the

attraction between protons (+) and electrons (-) increases as well, making each successive atom smaller than the previous one in THAT row (period).

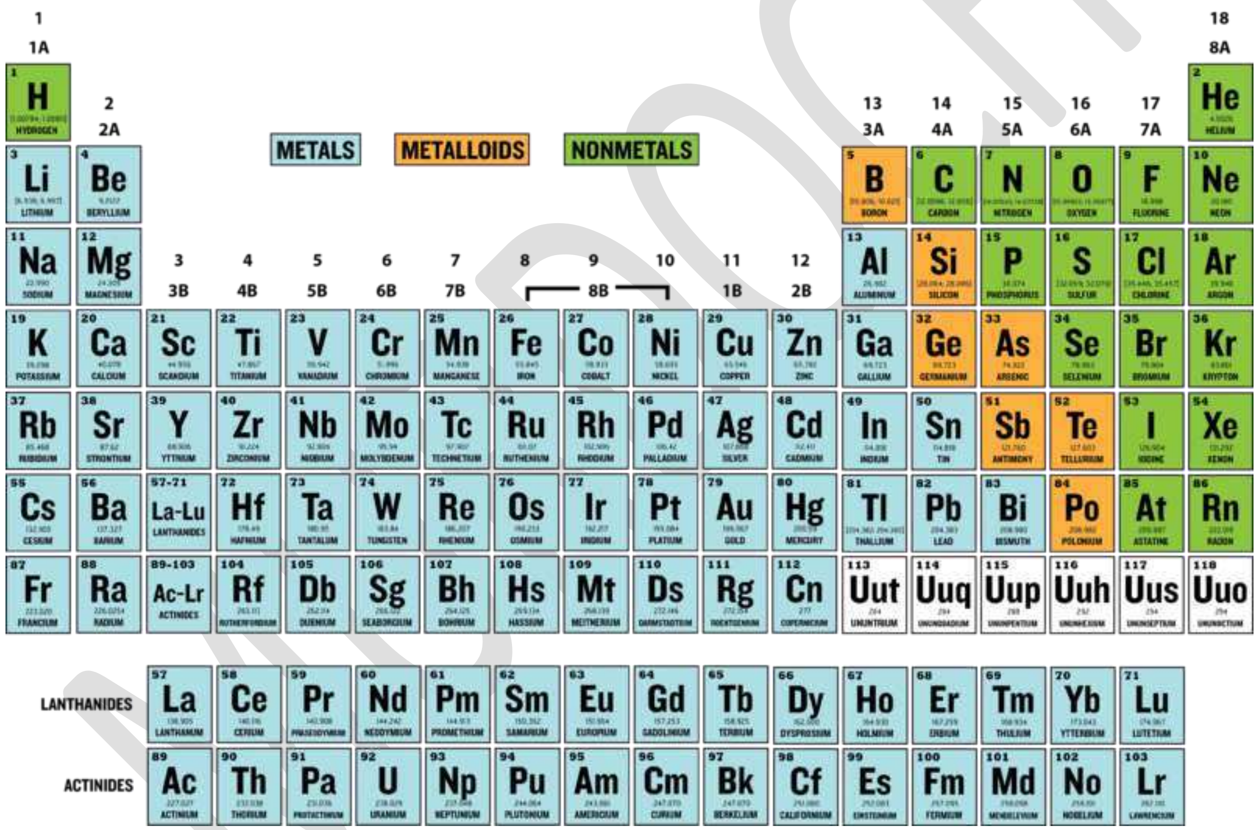
2. Within a **group** of the Periodic table, the atomic **radius** usually **increases** as the atomic number increases. This is due to an **additional** primary energy level (**PEL**) between the nucleus and the valence electrons, which increases the distance between the valence PEL and the nucleus. The more layers (PELs), the larger the atom.
   - As an example, for atoms in Group 1, Na (3 PELs) is larger than Li (2 PELs), and K (4 PELs) is larger than Na.



Topic: **Types of Elements**

Objective: How are elements on the Periodic Table arranged?

Elements on the Periodic Table are divided into three subgroups: metals, nonmetals, and metalloids (semimetals) as shown below.



These elements have distinct properties that give them distinct identities.

Topic: **Properties of Elements**

Objective: How are properties on the Periodic Table arranged?

Properties of Atoms as arranged on the Periodic table:

1. Electronegativity:
   a. Electronegativity is an atom's **attraction** for **electrons** in a chemical bond.
   b. Elements with a **small** atomic **radius** have a **greater** attraction for electrons, and therefore have a higher **electronegativity**. Electronegativity is measured on a relative scale, with fluorine having the highest electronegativity (4.0). Electronegativity may be found on Reference Table S.

2. First Ionization Energy:
   a. First Ionization Energy is the energy required to **remove** the most loosely held valence **electron** from the atom to **form** a **positive** ion when the atom is in the gas phase.
   b. First **ionization** energy is directly **proportional** to the **electronegativity**, because the more tightly an atom is attracted to its electrons, the more energy it is going to require for removing that electron.

c. In general, **metals** have **low** electronegativity and ionization **energies**, and tend to lose their valence electrons to form positive ions when bonding to nonmetal atoms. **Nonmetal** atoms have **high** electronegativity and ionization **energies**, and gain electrons from metal atoms to form negative ions, or bond with other nonmetals to form covalent bonds.

3. Metallic Character:

a. Metallic Character is the degree to which an element matches the characteristics of metals.

b. Metals **lose** electrons and form **positive** ions, therefore elements that have low electronegativity and easily lose electrons usually have high metallic character.

4. Nonmetallic character:

a. Nonmetallic character is the degree to which an element matches the characteristics of nonmetals. Nonmetals **gain** electrons and form **negative** ions, therefore elements which have high electronegativity and gain electrons easily have high nonmetallic character.

Topic: **Chemistry of Periodic Table**

Objective: What are the chemical properties of the Periodic Table?

Chemistry of Metals, Nonmetals, and Metalloids:

Type	EN & IN	Radius	What their ions do	Ion Charge	Properties
Metals	Low	Large	Lose (ionic)	Pos (+)	• Excellent conductors of heat and electricity • Malleable (may be hammered or rolled into thin sheets) • Ductile (may be drawn into thin wires) • Shiny (has luster) • Compose more than 2/3rds of the elements • Metallic character increases as ionization energy decreases. Francium is the most metallic element on the Periodic Table
Nonmetals	High	Small	Gain (ionic) Shared (covalent)	Neg (-)	• Poor conductors of heat and electricity • Brittle (shatters and/or crushes easily) • Dull appearance, not shiny like metals
Metalloids	Med.	Med.	Usually share	Either	• Semiconductors (sometimes conduct; sometimes not) • Used in making computer microchips • Has luster (like metals) and are Brittle (like nonmetals)

Topic: **Formation of Ions**

Objective: How are ions formed based on Periodic Table properties?

Formation of Ions:

- For every **electron** an atom **gains**, it becomes more **negatively** charged. If an atom gains three electrons when forming a bond, the atom becomes a -3 ion. For every **electron** an atom **loses**, it becomes more **positively** charged. If an atom loses two electrons when forming a bond, it becomes a +2 ion.
- When an atom becomes an ion, it does so by gaining or losing in such a way that the **ion** ends up having 8 valence electrons, or a Stable **Octet**, on the **outside** PEL of the ion.
- In forming an ion, the electrons are lost or gained from the valence 's' or 'p' sublevels. If an atom has **1** to **3** valence electrons, it wants to **lose** the electrons to form a stable valence octet.
  - Positive Ion:
    - Sodium (Na) has an atom electron configuration of 2-8-1. Sodium has one valence electron. Sodium could either gain 7 electrons (to make a stable octet 2-8-8), or lose 1 electron (to make a stable octet 2-8). Nature takes the easy (less energy) route, and sodium (2-8-1) does the latter and loses one electron to form a positive sodium ion (2-8).

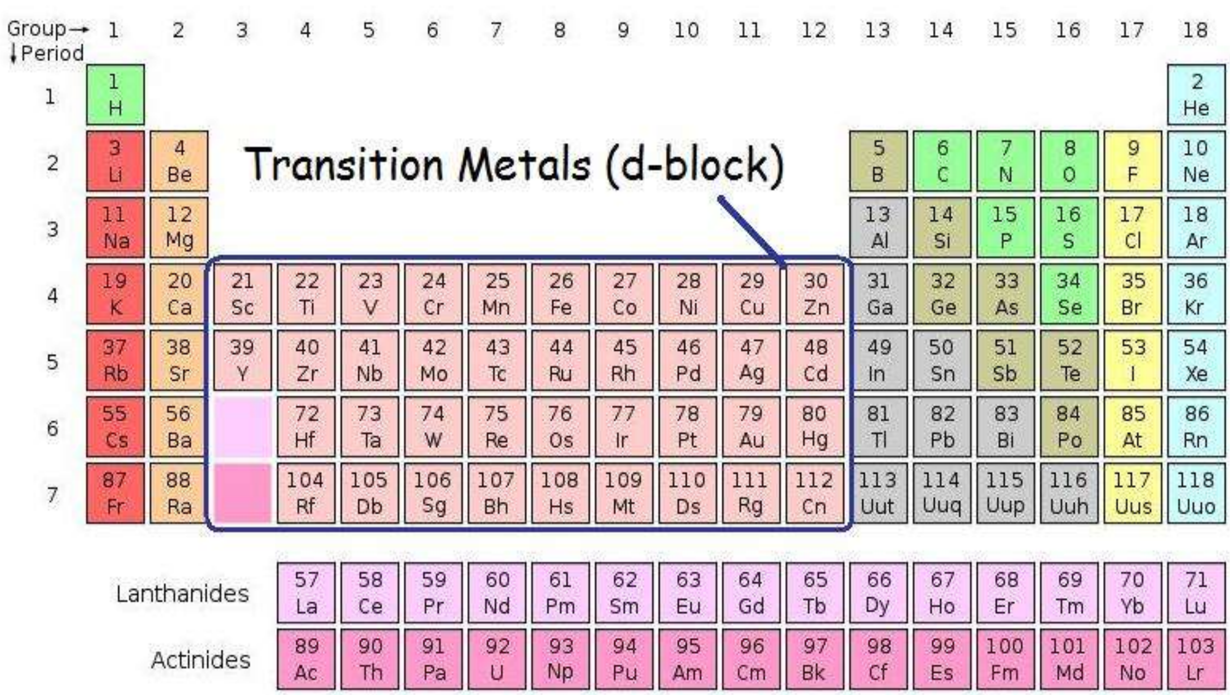
Topic: **Formation of Ions**

Objective: How are ions formed based on Periodic Table properties?

- If the atom forming an ion has **5** to **7** valence electrons, the atom wants to **gain** enough electrons to form a stable valence octet.
  - <u>Negative Ion</u>:
    - Chlorine (Cl) has an atom electron configuration of 2-8-7. Chlorine has seven valence electrons. Chlorine could either lose 7 electrons (to make a stable octet 2-8), or gain 1 electron (to form a stable octet 2-8-8). Again, nature takes the easy (less energy) route, and chlorine (2-8-7) does the latter and gains 1 electron becoming a negative chlorine ion (2-8-8).

Group	Valence Electron Configuration	# Valence Electrons	How it forms an ion	Charge of ion	Ion Valence Electron Configuration
1	…-1	1	loses 1	+1	…-8
2	…-2	2	loses 2	+2	…-8
13	…-3	3	loses 3	+3	…-8
14	…-4	4	loses 4 (to a more electronegative atom) gains 4 (from a less electronegative atom)	+4 -4	…-8
15	…-5	5	gains 3	-3	…-8
16	…-6	6	gains 2	-2	…-8
17	…-7	7	gains 1	-1	…-8
18	…-8	8	doesn't need to	0	…-8

You will note on the chart on the previous page that Groups 3 through 12 are missing. These elements are called Transition Elements, or sometimes **Transition** Metals. Elements in the middle of the Periodic Table (Groups 3 through 12) may **lose** electrons from both the **valence** energy level and the level below the valence level (first **kernel** level) when forming ions. This allows for a wide range of ion charges. Some of these elements may form more than one possible charge. Copper (Cu), for example, may form charges of +1 or +2, and iron (Fe) may form charges of +2 or +3, both of these examples depending on the circumstances of the reaction they are involved in.

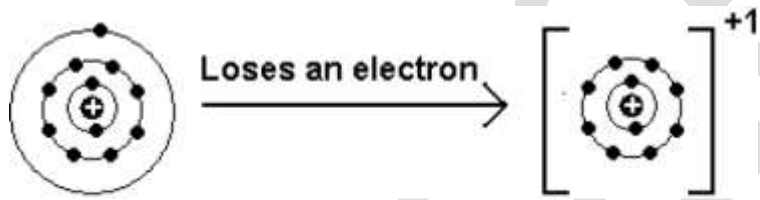


Topic: **Ionic Radius**

Objective: How does the overall size of ions change versus atoms?

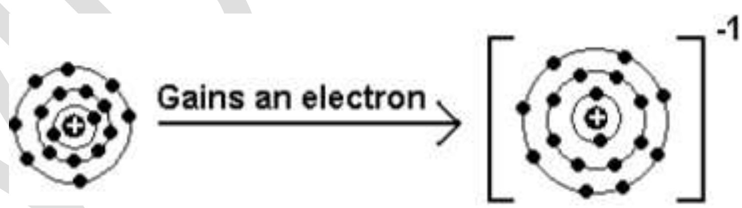
Ionic Radius:

- When an atom **GAINS** electron(s), its radius **increases**. When an atom **LOSES** electron(s), its radius **decreases**.



Sodium Atom Sodium Ion

i. Sodium starts with 3 PELs (2-8-1) and loses its one valence electron with a new electron configuration of (2-8). Sodium now has 2 PELs and the NA^{+1} ion has a smaller radius than Na^{0}. Metal ions have a SMALLER radius than metal atoms.



Chlorine Atom Chlorine Ion

ii. Chlorine has 3 PELs (2-8-7) and gains one electron to form a stable octet. The additional electron makes Cl^{-1} larger than Cl^{0} as the additional electron increases the repulsive force between the valence electrons. Nonmetal ions have a LARGER radius than nonmetal atoms.

Topic: **Naming of Ions**

Objective: How are Ions named based on the original Atoms?

Naming of Ions:

The names of ions are dependent on the charge of the ion.

1. Positive Ions keep the SAME **name** as the **element**. If the atom is capable of forming more than one possible ion, a Roman **numeral** is placed after the ion name, **signifying** the ionic **charge**.

Ion	Name	Ion	Name	Ion	Name
Na^{+1}	sodium	Fe^{+2}	iron (II)	Pb^{+2}	lead (II)
K^{+1}	potassium	Fe^{+3}	iron (III)	Pb^{+4}	lead (IV)
Ca^{+2}	calcium	Cu^{+1}	copper (I)	Cr^{+2}	chromium (II)
Mg^{+2}	magnesium	Cu^{+2}	copper (II)	Cr^{+3}	chromium (III)
Ag^{+1}	silver	Au^{+1}	gold (I)	Sn^{+2}	tin (II)
Al^{+3}	aluminum	Au^{+3}	gold (III)	Sn^{+4}	tin (IV)

- The use of Roman numerals to identify ionic charge is called the Stock System.

2. Negative ions are named after the element, with the element's second syllable replaced with the suffix "**-ide**".

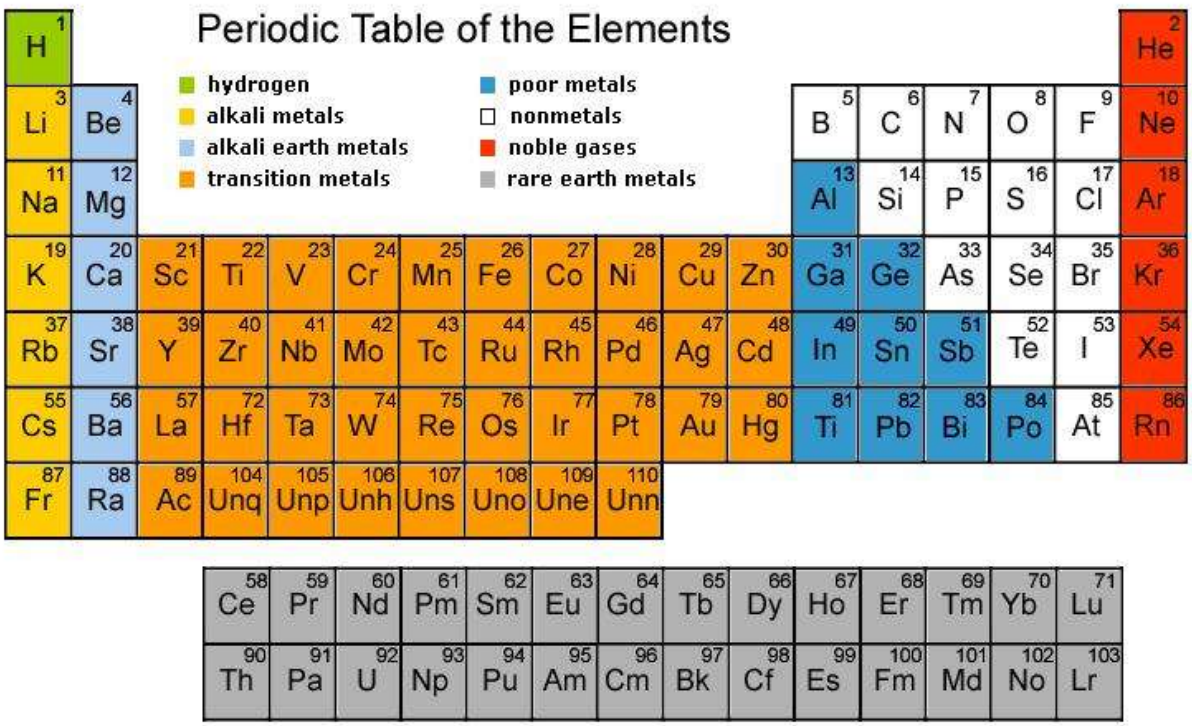
Ion	Element Name	Ion Name	Ion	Element Name	Ion Name
O^{-2}	oxygen	Oxide	S^{-2}	Sulfur	sulfide
N^{-3}	nitrogen	Nitride	P^{-3}	Phosphorous	phosphide
H^{-1}	hydrogen	Hydride	Cl^{-1}	Chlorine	chloride

Topic: **Chemistry of Periodic Groups**

Objective: How are chemical properties based on the periodic table?

Chemistry of the Groups of the Periodic Table:

The groups (**vertical** columns) of the Periodic Table are grouped according to **similar** chemical **properties**.



Group	Name	# of e^- in Valence	Ion charge	Notes
1	Alkali Metals	1	+1	• Extremely reactive; only found in compounds • May be extracted from compounds using electricity • Reacts violently with water forming $H_{2(g)}$ and a base • Alkali means base (as opposed to acid)
2	Alkaline Earth Metals	2	+2	• Very reactive, only found in compounds • Can be extracted from compounds using chemical reactions • Reacts quickly with water to form $H_{2(g)}$ and a base • Alkaline means base
3-12	Transition Metals	Varies	+1 to +7	• Range of reactivity; some quite reactive, others nonreactive • Some can be found in pure form in nature; Cu, Ag, Au, etc. • Ions are colored, so compounds with transition elements in them are often colored • Many form multiple charges; Stock System to name those ions
17	Halogens	7	-1	• Extremely reactive and corrosive, only found in compounds • May be extracted from compounds using electricity • Reacts violently with metals to form halide compounds (NaCl)
18	Noble Gases	8	0	• Nonreactive; not found in compounds • Xe and Kr may be forced to react with F_2 in a lab • Have a stable valence octet, and they have no need to bond

Phases on the Periodic Table:

- All elements on the Periodic Table are in the solid phase at 25°C except:
  - Mercury & Bromine – liquid
  - N, O, F, Cl, H, He, Ne, Ar, Kr, Xe, & Rn – gases

Topic: **Elemental Molecules**

Objective: What are the elemental (same element) molecules?

Elemental molecules:

- Molecule: A particle made of nonmetal atoms that are covalently bonded together.

1. Monoatomic molecules:
   i. Noble gases: **Noble** gases do not react with other elements, so the individual atoms of a noble gas (He, Ne, Ar, Kr, Xe, & Rn) are considered to be "**monoatomic** molecules" of that gas.
2. Diatomic molecules:
   i. Reactive nonmetals: Two individual atoms of the **SAME** element that are reactive may bond together to form a **diatomic** molecule. These elements are found normally in the diatomic state unless bonded with a different type of atom.
      - Br_2, I_2, N_2, Cl_2, H_2, O_2, & F_2

Student name: ________**Key**_________ *Class Period:* _*3, 5, 10*_

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The Periodic Table homework

1. What do all elements in a period have in common?
   - ***a) Number of PELs***
   - b) Number of valence electrons
   - c) Number of electrons
   - d) Number of protons

2. What do all elements in Group 2 have in common?
   - a) Number of PELs
   - ***b) Number of valence electrons***
   - c) Number of electrons
   - d) Number of protons

3. Using Reference Table S and the Periodic Table, what change occurs to the atomic radius of the atoms of Na through Cl from left to right?
   - a) Increases
   - ***b) Decreases***
   - c) Remains constant

4. Using Reference Table S and the Periodic Table, what change occurs to the atomic radius of the elements in Group 2 from top to bottom?
   - ***a) Increases***
   - b) Decreases
   - c) Remains constant

5. Why is the radius of bromine larger than the radius of fluorine?
   - a) More electrons
   - ***b) More PELs***
   - c) More nuclear charge
   - d) More neutrons

6. Why is the radius of oxygen smaller than the radius of boron?
   - a) More electrons
   - b) More PELs
   - ***c) More nuclear charge***
   - d) More neutrons

7. How do ions of metal atoms form?
   - a) Gain electrons
   - ***b) Lose electrons***
   - c) Gain protons
   - d) Lose protons

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8. As you move left to right in period 5, what happens to the degree of metallic character of an element?
   a) Increases
   b) Decreases
   c) Remains constant

9. As you move bottom to top in group 14, what happens to the degree of nonmetallic character of an element?
   a) Increases
   b) Decreases
   c) Remains constant

10. How do ions of nonmetal atoms form?
    a) Gain electrons
    b) Lose electrons
    c) Gain protons
    d) Lose protons

11. How many valence electrons do negative ions have?
    a) 1 b) 2 c) 6 ***d) 8***

12. What changes occur to the radius of an atom when it becomes a + ion?
    a) Increases
    b) Decreases
    c) Remains constant

13. What changes occur to the radius of an atom when it becomes a – ion?
    a) Increases
    b) Decreases
    c) Remains constant

14. Which of the following elements exist as monoatomic molecules at STP?
    a) Li b) Br ***c) Ne*** d) Au

15. Which of the following elements exists as diatomic molecules at STP?
    a) Li ***b) Br*** c) Ne d) Au

*Student name: ________ **Key** ________* *Class Period: _3, 5, 10_*

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<u>The Periodic Table homework</u> (Cont'd)

16. Name the following ions *or* write the symbol and charge.

Ion	Name	Ion	Name	Ion	Name
Na^{+1}	sodium ion	Au^{+3}	gold(III) ion	O^{-2}	oxide
Br^{-1}	bromide ion	Mg^{+2}	magnesium ion	Pb^{+4}	lead (IV)
Zn^{+2}	zinc ion	Cu^{+1}	copper(I) ion	Cr^{+3}	chromium (III)

17. Using Reference Table S, determine which element in the following pairs is larger.

Compare… pm	Which has the larger radius?	Compare… pm	Which has the larger radius?
N_{71} or O_{64}	N	Cl_{100} or I_{136}	I
N_{71} or P_{109}	P	Mg_{140} or Na_{160}	Na
O_{64} or S_{104}	S	Mg_{140} or P_{109}	Mg
P_{109} or S_{104}	P	Li_{130} or Be_{99}	Li
P_{109} or As_{120}	As	Ca_{174} or Ba_{206}	Ba

18. Using Reference Table S, determine which atom in the following pairs has the higher electronegativity (EN) and higher ionization energy (IE).

Compare…	Higher EN	Higher IE kJ/mol	Compare…	Higher EN	Higher IE kJ/mol
N or O	O 3.4	N 1402	Cl or I	Cl 3.2	Cl 1251
N or P	N 3.0	N 1402	Mg or Na	Mg 1.3	Mg 738
O or S	O 3.4	O 1314	Mg or P	P 2.2	P 1012
P or S	S 2.6	P 1012	Li or Be	Be 1.6	Be 900
P or As	Same	P 1012	Ca or Ba	Ca 1.0	Ca 590

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19. Using the Period Table, determine the number of valence electrons in atoms of the following elements along with the Principal Energy Level of the valence e^-.

Element	# Valence Electrons	PEL	Element	# Valence Electrons	PEL
Li	1	2	Na	1	3
Mg	2	3	Ca	2	4
Al	3	3	Ga	3	4
Ge	4	4	Sn	4	5
N	5	2	P	5	3
Se	6	4	Te	6	5
Cl	7	3	I	7	5
Kr	8	4	Rn	8	6

20. Determine the charge for the ions of each of the elements below and indicate the change in the ionic radius compared to the original atom.

Ion	Charge	Ion larger or smaller than original atom?	Ion	Charge	Ion larger or smaller than original atom?
Li	+1	S	Mg	+2	S
Na	+1	S	Ca	+2	S
K	+1	S	Sc	+3	S
H	+1	S	N	-3	L
P	-3	L	O	-2	L
S	-2	L	F	-1	L
Cl	-1	L	Br	-1	L

Notes page:

Topic: **Ionic Bonding**

Objective: Why do atoms join to form molecules and compounds?

Chemical Bond: A chemical **bond** results from the **competition** for valence **electrons** between two atoms. Chemical bonds are what hold atoms together to form compounds, and chemical bonds are broken when compounds are decomposed back into the original elements.

[Watch Crash Course Chemistry Chemical Bonds video](https://www.youtube.com/watch?v=QXT4OVM4vXI)

https://www.youtube.com/watch?v=QXT4OVM4vXI

Ionic Bonding:

- Ionic bonding occurs between a **metal** atom and a **nonmetal** atom.
- The **nonmetal** atom has a higher **electronegativity** than the metal atom, and therefore '**wins**' the competition for each atom's valence electrons. The nonmetal atom gains electrons from the metal atom, as the metal atom loses ALL of its valence electrons to the nonmetal.
- The number of electrons **lost** or **gained** will be the number needed by each atom to **form** a stable **octet** of eight valence electrons.
- The **metal** atom **loses** electrons (oxidation) and forms a + charged cation.

- The **nonmetal** atom **gains** electrons (reduction) and forms a -charged anion.
- The now oppositely charged **cation** and **anion** attract each other. This **charge** attraction **forms** the ionic **bond**. Ionic attraction is surface attraction, and is easily broken by melting or dissolving in water.

Determining Ionic Bonding:

- To determine if a bond is ionic, look at the electronegativity of both bonding elements in Reference Table S and subtract lower from higher. If the **electronegativity** difference is above **1.7** the atom with the higher electronegativity has **enough** attraction to **remove** the **electrons** from the atom with the lower electronegativity.

Topic: **Forming Ionic Bonds**

Objective: How does and ionic bond form?

<u>How does an ionic bond form</u>?

1. Formation of an ionic bond between sodium and chlorine.

   Element types? Na (METAL) Cl (NONMETAL)

   EN = electronegativity END = electronegativity difference

   EN of Na = 0.9 EN of Cl = 3.2 END = 2.3 (so 2.3 > 1.7)

   i. The sodium atom loses its one valence electron (oxidation) to form a Na^{+1} cation.
   ii. The chlorine atom gains (reduction) the one electron sodium lost from its valence and becomes a Cl^{-1} anion.
   iii. The Na^{+1} and Cl^{-1} ions, now oppositely charged, are attracted and form ionic bonded NaCl (sodium chloride).
   iv. This bond may easily be broken by heating NaCl to the melting point as that provides enough energy for the ions to separate. Dissolving NaCl in water will also break the ionic bond, with the now separate Na^{+1} and Cl^{-1} ions clinging to the water molecules.

Topic: **Ionic Compound Properties**

Objective: What properties do Ionic Compounds share?

Properties of Ionic Compounds:

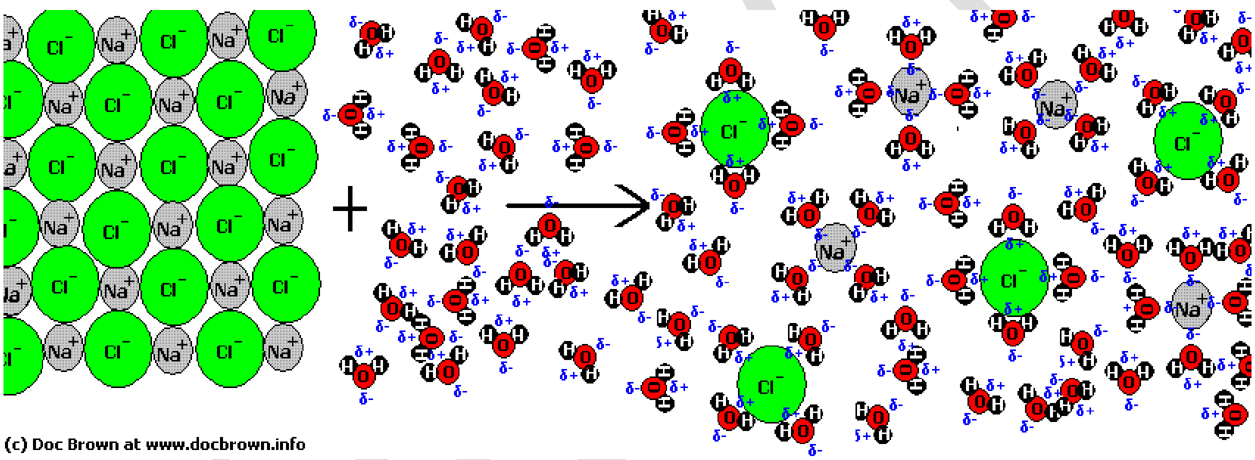
1. Ionic solids have high melting and boiling points.
2. Ionic liquids have low vapor pressures; i.e. they don't evaporate easily.
3. Ionic solids are brittle, meaning they crush easily into powder.
4. Ionic **liquids** and **solutions** conduct electricity as the charged particles are free to move around and **carry** their electrical **charge** from one area to another area. Ionic solutions (like salt water) are called electrolytes because of this ability to conduct electricity.
5. Ionic **solids** do **NOT** conduct electricity as the ions are held tightly together in a crystal lattice and the ions do not move.

Topic: **Ionic Crystal Structure**

Objective: What properties do Ionic Crystals have?

Ionic Crystal Structure:

- In an ionic solid, **melting** or **dissolving** in water breaks the crystals into an ionic liquid or aqueous solution composed of **free**-moving ions.



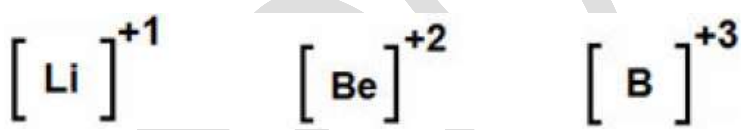
- Note that in the **SOLID** IONIC CRYSTAL (shown above left), the ions are **locked** in place, unable to move. The locked (bonded) ions cannot conduct electricity. If the ionic bonds are broken by **melting** or **dissolving**, the **ions** are now **free** and may **conduct** electricity (shown above right).

Topic: **Ionic Bonding Diagrams**

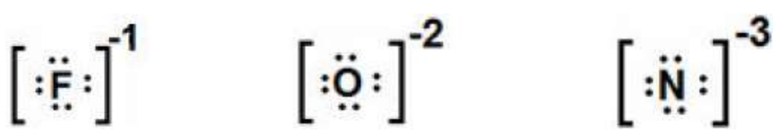
Objective: How do we represent Ionic Bonding with Lewis Diagrams?

Ionic Bonding Dot Diagrams:

1. When a **metal** atom **loses** its valence electron(s), the oxidized atom now has zero electrons in what ***was*** its valence PEL. To write the dot diagram of the metal ion, write the atomic symbol, place brackets around the symbol, and write the ***positive*** ionic charge outside of the brackets at the upper right.



2. When a **nonmetal** atom **gains** valence electrons, the reduced atom now has eight electrons in its valence PEL. To write the dot diagram of the nonmetal ion, write the atomic symbol, place the eight dots around the symbol, place brackets around the dots, and write the ***negative*** ionic charge outside of the brackets at the upper right.



Topic: **Ionic Bonding Diagrams**

Objective: How do we represent Ionic Bonding with Lewis Diagrams?

3. To draw the dot diagram of the ionic compound, place the dot diagrams next to each other showing:
   i. The ion charges **cancel** each other (add up to **zero**);
   ii. The **opposite** charged ions are **near** to each other, and the **like** charged ions are as **far** away from each other as possible.

<u>Example Ionic Bonding Dot Diagrams</u>:

Formula	Dot Diagram
LiF	$[\mathrm{Li}]^{+1}\ [:\ddot{\underset{..}{\mathrm{F}}}:]^{-1}$
BeO	$[\mathrm{Be}]^{+2}\ [:\ddot{\underset{..}{\mathrm{O}}}:]^{-2}$
Li_2O	$[\mathrm{Li}]^{+1}\ [:\ddot{\underset{..}{\mathrm{O}}}:]^{-2}\ [\mathrm{Li}]^{+1}$
BeF_2	$[:\ddot{\underset{..}{\mathrm{F}}}:]^{-1}\ [\mathrm{Be}]^{+2}\ [:\ddot{\underset{..}{\mathrm{F}}}:]^{-1}$

Topic: **Covalent Bonding**

Objective: How do two nonmetal atoms bond without sharing?

Covalent Bonding:

- Covalent Bonding involves **sharing** of **electrons**, not give ***or*** take.
- If two **NONMETAL** atoms attempt to gain each other's valence electrons, they do not have enough difference in electronegativity to 'give' or 'take', so the nonmetals **share** electrons.
- The electrons **shared** are the **unpaired** valence electrons.
- The bonded **electrons** actually become a **part** of each **atom**. This makes a covalent bond much STRONGER than an ionic bond. Covalent bonds may not be broken by melting, or dissolving in water, so covalent compounds do not conduct electricity readily, regardless of phase.

Topic:	**Covalent Bonding**
Objective: How do two nonmetal atoms bond without sharing?	

Determining Covalent Bonding:

- To determine if a bond is covalent, check the electronegativity difference using Reference Table S. If the **electronegativity** difference is **less** than **1.7**, the atom with the higher electronegativity is not able to pull electrons from the atom with the lower electronegativity. This forces the atoms to share unpaired valence electrons, meaning the **bond** is **covalent**.
- Each covalently bonded atom in the compound will now have eight valence electrons, excepting hydrogen, which has only a 1s sublevel holding a maximum of two electrons in its valence PEL.

How many Covalent Bonds may a nonmetal atom form?

Nonmetal	Dot Diagram	# unpaired e-	# of covalent bonds	Nonmetal	Dot Diagram	# unpaired e-	# of covalent bonds
N	·N̈· (with one dot below)	3	3	S	:S̈· (with one dot below)	2	2
O	:Ö· (with one dot below)	2	2	Cl	:C̈l· (with two dots below)	1	1
F	:F̈· (with two dots below)	1	1	P	·P̈· (with one dot below)	3	3
C	·Ċ· (with one dot below)	4	4	Br	:B̈r· (with two dots below)	1	1
H	Ḣ	1	1	I	:Ï· (with two dots below)	1	1

Topic: **Forming Covalent Molecules**

Objective: How do two nonmetal atoms form covalent bonds?

Forming Covalently Bonded Molecules:

- Molecules are formed from **nonmetal** atoms covalently **bonding** together. Each **molecule** of the same substance has a unique molecular formula that tells you exactly how many atoms of each element are found in the molecule.

i. Water (H_2O) is a molecule made of two hydrogen (H) atoms covalently bonded to one oxygen (O) atom.

ii. Methane (CH_4) is a molecule made of one carbon (C) atom covalently bonded to four hydrogen (H) atoms.

iii. Ammonia (NH_3) is a molecule made of one nitrogen (N) atom covalently bonded to three hydrogen (H) atoms.

Types of Covalent Bonding:

There are two types of covalent bonding; Polar Covalent and Nonpolar Covalent bonding

Watch Crash Course Chemistry Polar & Nonpolar Molecules video

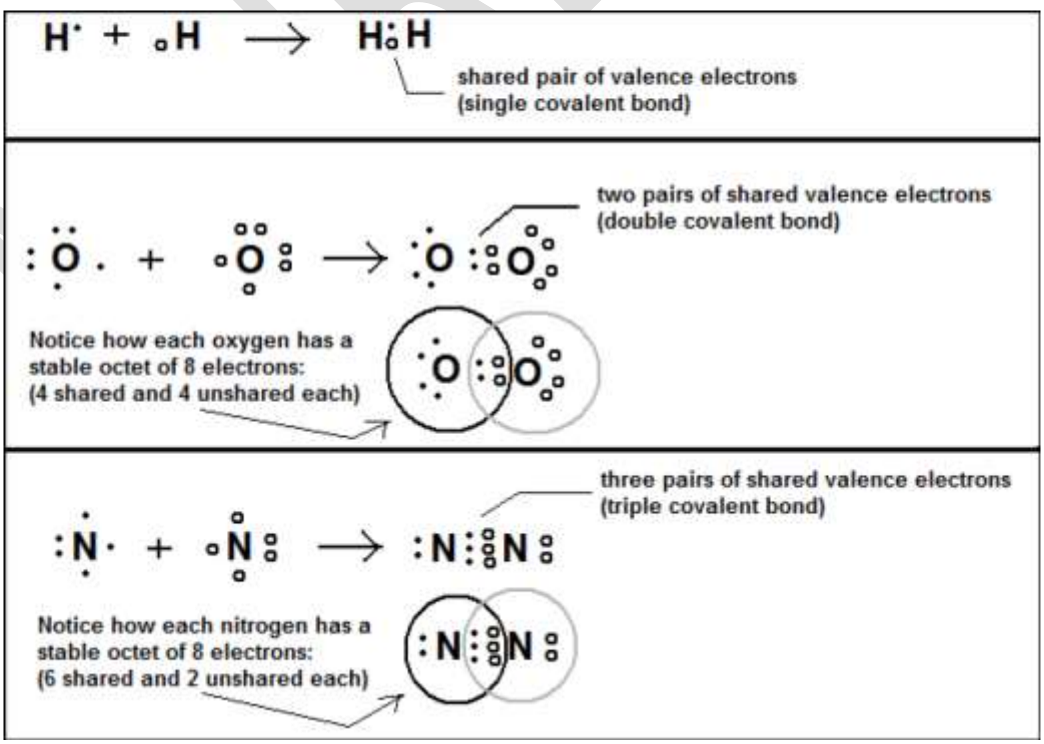
https://www.youtube.com/watch?v=PVL24HAesnc

Topic: **Nonpolar Covalent Bonding**

Objective: How does low electronegativity difference change bonds?

Nonpolar Covalent Bonds:

- Nonpolar Covalent Bonds are formed between nonmetal atoms that have a **difference** in **electronegativity** between 0 and 0.4.
- In a **nonpolar** covalent bond the **electrons** are **equally** shared between the atoms. Examples of nonpolar covalent bonding include the diatomic molecules (Br_2, I_2, N_2, Cl_2, H_2, O_2, & F_2), which are formed when not in the presence of other nonmetal or metal atoms. Each covalent diatomic molecule has a stable octet of valence electrons.



Topic: **Polar Covalent Bonding**

Objective: Does medium electronegativity difference change bonds?

Polar Covalent Bonds

- Polar Covalent Bonds are formed between nonmetal atoms with a **difference** in **electronegativity** between **0.5** to around **1.7**.
- In a **polar** covalent Bond the **electrons** are shared **unequally** between the atoms within the bond. The electrons spend more time near the more electronegative atom, giving the more electronegative atom a partial negative charge. Therefore the less electronegative atom(s) become partially positive charged. The **charged** ends of the bonds form **POLES** (opposite charged ends) which is why the bond(s) are called "polar".
- To denote the partiality of the charge, we use the lower-case Greek letter "delta": δ (δ$^+$ = **partially** positive; δ$^-$ = **partially** negative)

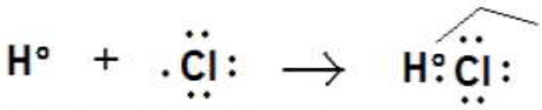
Bonding Nonmetal Atoms	Electronegativity Of Each Atom (END)	Which Pole Is δ+ and which is δ-?
H and Cl	H: 2.2 Cl: 3.2 (1.0)	δ+ H-Cl δ-
H and O	H: 2.2 O: 3.5 (1.3)	δ+ H-O δ-
O and N	O: 3.5 N: 3.0 (0.5)	δ- O-N δ+
C and F	C: 2.6 F: 4.0 (1.4)	δ+ C-F δ-

Topic: **Covalent Molecular Bonding**

Objective: How do atoms Covalently Bond to form Molecules?

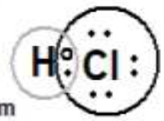
Covalently Bonded Atoms in Molecules:

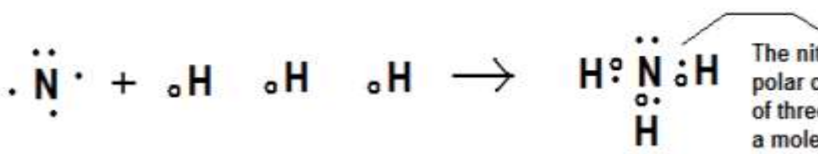
- The diagrams on the next page show how each bonding atom's unpaired electrons pair up, with the atoms then becoming a part of one another. Note that this is different than ionic bonding, where electrons are transferred from one atom to another. In a covalent bond, the valence electrons of one bonding atom are shown as Lewis dots, while the valence electrons of another bonding atom are shown as small circles. This is done to clearly see which atom the shared electrons in the bond are originally from.
- The particle formed by the **covalent** bonding atoms is called a **MOLECULE**. Molecules are made of nonmetal atoms bonding with other nonmetal atoms. There are some exceptions, but for Regents Chemistry molecules will be considered made of only nonmetals.
- When the atoms bond and share their unpaired valence electrons from their individual atomic orbitals, the newly **shared** pair of **electrons** forms a molecular **orbital** belonging to **both** atoms. This makes the covalent bond MUCH stronger and more difficult to break than ionic bonds. Covalent bonds cannot usually be broken by melting or dissolving.



The chlorine has formed a single polar covalent bond with a hydrogen atom. This forms a molecule of HYDROGEN CHLORIDE, HCl. When dissolved in water, it forms HYDROCHLORIC ACID, HCl(aq).

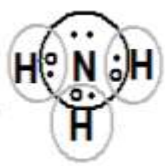
Notice how the chlorine has eight valence electrons, two that are shared with the hydrogen, and 3 unshared pairs of its own (shown here on the top, right side and bottom of the Cl atom). The hydrogen has 2, which is all it needs.

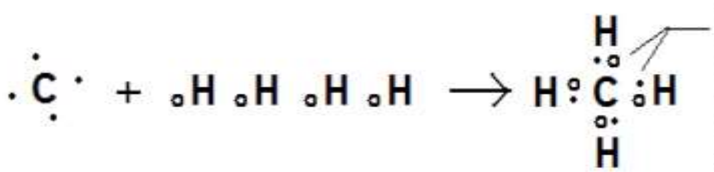




The nitrogen has formed three single polar covalent bonds, one with each of three hydrogen atoms. This forms a molecule of AMMONIA, NH_3.

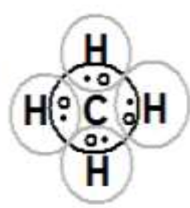
Notice how the nitrogen has eight valence electrons, two that are unshared (the ones on top) and three pairs of shared electrons. Each hydrogen now has two valence electrons (all from shared pairs), which is all the tiny structure of hydrogen can handle.





The carbon has formed four single nonpolar covalent bonds, one with each of four hydrogen atoms. This forms a molecule of METHANE, CH_4. Why nonpolar? C has an EN of 2.6 and H is 2.2, which is an END of 0.4. Between 0 and 0.4 is NONPOLAR.

Notice how the carbon has eight valence electrons, all from four shared pairs. Each hydrogen now has two valence electrons, which is all it needs.



[Watch Crash Course Bonding Models and Lewis Structures video](https://www.youtube.com/watch?v=a8LF7JEb0IA)

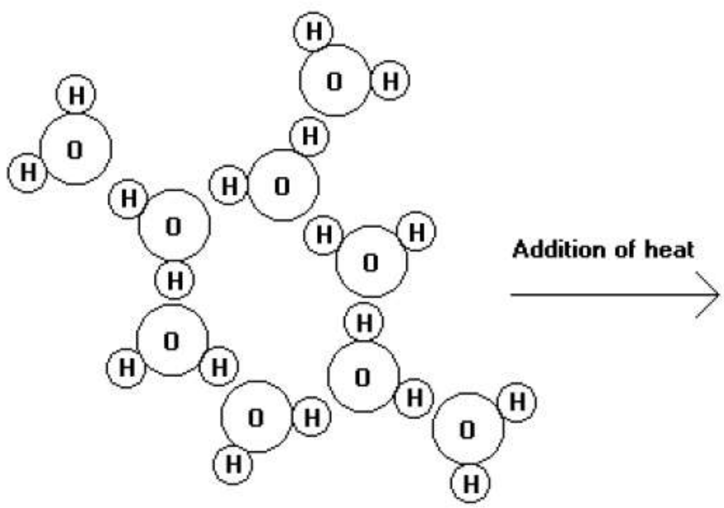
https://www.youtube.com/watch?v=a8LF7JEb0IA

Topic: **Covalent Molecular Lattice**

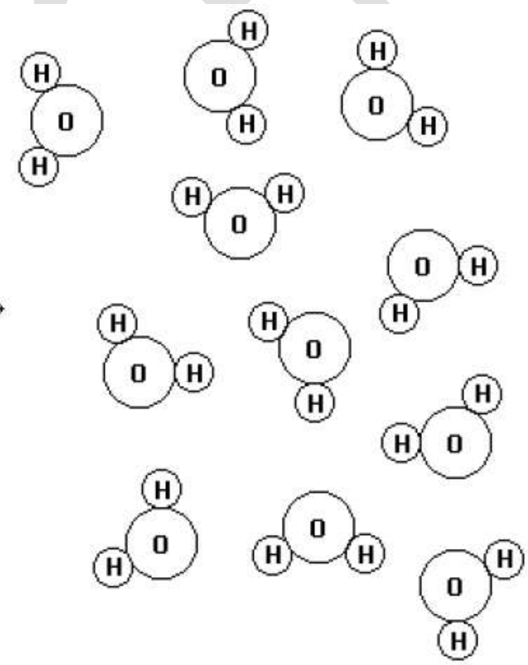
Objective: How do atoms Covalently Bond to form Lattice structures?

<u>Covalent (Molecular) Crystal Lattice</u>:

Covalent (Molecular) crystal lattice may melt (or be dissolved) to form freely-moving molecules.



Covalent Crystal Lattice: molecules of water are attracted to each other with InterMolecular Attractive Forces. These hold the molecules together in a regular geometric pattern. Since IMAF are weaker than ionic attractions, these crystal lattices tend to be weak, and therefore soft and easy to melt.



When melted, the molecules separate from each other. The covalent bonds do not break, just the IMAF. The molecular liquid does not have charged particles like ionic liquids do, therefore molecular liquids cannot conduct electricity.

Notice in the solid molecular crystal the molecules are locked in place, unable to move freely. There are **no** charged **ions**; therefore the liquid (or solution) molecules do **not** carry current. The exceptions are acids (more on this later…☺).

Topic: **Bonding Summary**

Objective: How do the bonding types compare to each other?

Bonding Type Summary:

END=Electronegative Difference EN=Electronegativity M=Metal NM=Nonmetal

Bond Type	Bonded Elements	END	Substance name	How the bond forms
Ionic	M-NM	> 1.7	Ionic	• Metal atom (low EN) loses an e^- to nonmetal atom (high EN) • Metal atom is OXIDIZED; nonmetal atom is REDUCED • Metal atom forms + ion; nonmetal atom forms - ion • Oppositely charged (+ cation & - anion) ions attract • The attraction between the ions IS the ionic bond
Covalent	NM-NM	0 - 1.7	Molecular	NONPOLAR COVALENT • Two nonpolar atoms with an END of 0 - 0.4 share their unpaired valence electrons EQUALLY • No oppositely charged ends of atoms POLAR COVALENT • Two nonmetal atoms with an END of 0.5 or higher share their unpaired valence electrons UNEVENLY • The atom with the lower electronegativity develops a partially positive charge (δ^+) • The atom with the higher electronegativity develops a partially negative charge (δ^-)

Topic: **Bonding Reminders**

Objective: How do the bonding types compare to each other?

Bonding Type Reminders:

Ionic Substances	• Forms ionic crystal lattices with very high melting and boiling points • Electricity may be conducted by charged particles (ions) • The ionic bond breaks when the compound is melted or dissolved in water • Breaking the bonds forms ions that are capable of flowing • Ionic liquids and aqueous solutions are good conductors of electricity • Ionic solids cannot conduct electricity, since the ions are locked in a crystal lattice and are not free to move
Covalent Substances	• Molecules are particles made of covalently bonded nonmetal atoms • As no ions are formed, molecular substances do not normally conduct electricity • Molecules may have partially charged ends, like a magnet, due to polar covalent bonds • Partially charged ends of molecules have much lower charge than ions, and therefore have low melting and boiling points • The attraction of one molecule's δ^{+} end towards the δ^{-} end of another molecule is called an Intermolecular Attractive Force (IMAF) • IMAF is what causes water to have surface tension (cohesion between water molecules and adhesion between water and other surfaces) • Atomic orbitals in each bonded atom combine in such a way that the shared electrons belong to BOTH of the bonded ATOMS and form molecular orbitals

Student name: _________**Key**_________ *Class Period:* _***3, 5, 10***_

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Bonding homework

IONIC BONDING:

1. Which of the following compounds is formed by ionic bonding?
   a) C_2H_6 b) NO_2 ***c) Li_2O*** d) O_2

2. Which molecule below will conduct electricity (is an electrolyte)?
   a) $NaCl_{(s)}$ b) $N_{2(s)}$ ***c) $LiF_{(aq)}$*** d) $CaI_{2(s)}$

Circle the answer below when potassium and chlorine atoms bond together.

3. Potassium (gains/***loses***) _1_ (#) of valence electron(s) and becomes a (**+** / -) ion, called a (***cation***/anion) during (***oxidation***/reduction).
4. Chlorine (***gains***/loses) _1_ (#) of valence electron(s) and becomes a (+ / **-**) ion, called a (cation/***anion***) during (oxidation/***reduction***).
5. Draw the dot diagram of the ionic compound KCl.

COVALENT BONDING:

6. Which of the molecules listed below has the most polar bond between the bonded atoms?
   a) HF b) HCl c) HBr d) HI

7. Which of the following compounds is formed by covalent bonding?
   a) Na_2S b) $AlCl_3$ ***c) $C_6H_{12}O_6$*** d) LiH

8. Which of the following molecule contains a nonpolar covalent bond?
   a) H_2O b) HF ***c) F_2*** d) NH_3

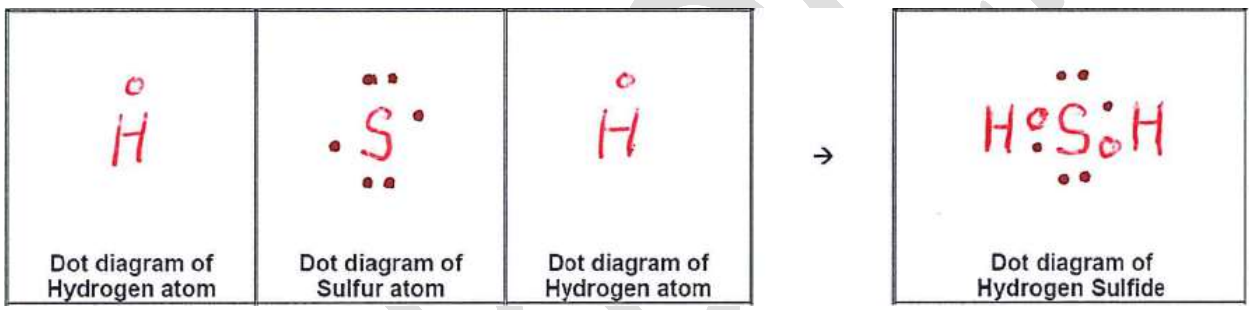
Cont'd next page

Circle the answer below when an atom of hydrogen and an atom of fluorine bond together.

9. Hydrogen will be partially (***positive***/negative) charged because it has a (higher/***lower***) electronegativity than fluorine.
10. Fluorine will be partially (positive/***negative***) charged because it has a (***higher***/lower) electronegativity than hydrogen.

Hydrogen and sulfur atoms combine to form hydrogen sulfide (H_2S).

11. Using the boxes below, show how hydrogen and sulfur combine to form a molecule of hydrogen sulfide.



Can you tell if the formation of H_2S is covalent or ionic bonding? Explain using electronegativity difference.

Covalent, as H has an EN of 2.2 and S has an EN of 2.6, and ionic bonding usually requires an END of 1.7 or more.

Are the bonds between hydrogen and sulfur polar or nonpolar? Explain using electronegativity difference.

H has an EN of 2.2 and S has an EN of 2.6, and polar covalent bonds usually require an END of 0.5 or more.

Student name: _________**Key**_________ *Class Period:* _**3, 5, 10**_

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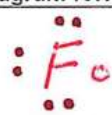
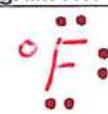
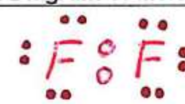
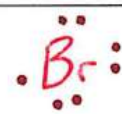
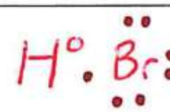
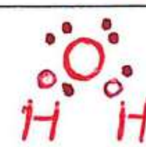
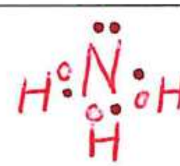
Bonding homework (Cont'd)

Complete the following chart by drawing the dot diagram of each element in the molecule and then the dot diagram of the molecule together. If the formula has more than two atoms (like H_2O), make sure you show all the atoms you have.

Formula	Dot diagram for:	Dot diagram for:	Dot Diagram of Molecule
F_2	F 	F 	
N_2	N 	N 	
HBr	H 	Br 	
H_2O	H 	O 	
NH_3	N 	H 	

Identify the following bonds as being polar covalent or nonpolar covalent. For the polar covalent bonds, label the δ^+ and δ^- ends.

Bond EN EN	END	Polar or Nonpolar?	If polar, label the δ^+ and δ^- ends
2.2 H – H 2.2	0	Np	H – H
2.2 H – C 2.6	0.4	Np	H – C
2.2 H – Cl 3.2	1.0	P	δ^+H – Cl δ^-

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Notes page: