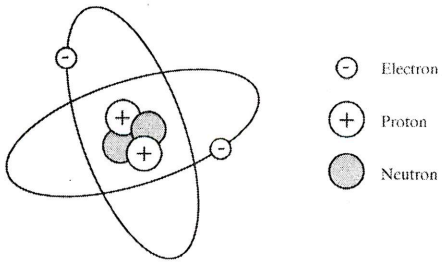
Name: KEY Date: _____ Block: _____

CHEMISTRY NOTES PACKAGE (2020-2021: SCIENCE 10)

Subatomic Particles Calculations

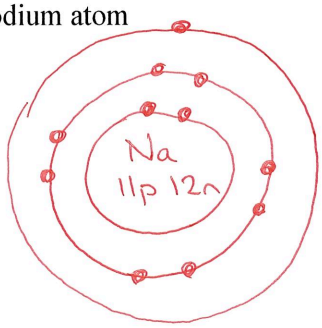
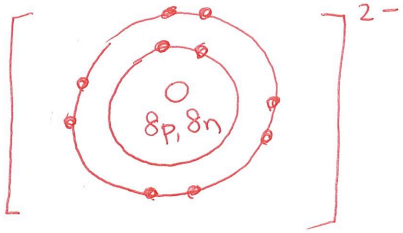
	Protons (p)	Neutrons (n)	Electrons (e)
Atom (neutral)	atomic number	atomic number minus rounded atomic mass	atomic number
Ion (charged)			atomic number minus ion charge



	p	n	e		p	n	e
Na atom	11	12	11	**O atom**	8	8	8
Na^+ ion	11	12	10	**O^{2-} ion**	8	8	10
Mg atom	12	12	12	**Cl atom**	17	19	17
Mg^{2+} ion	12	12	10	**Cl^- ion**	17	19	18

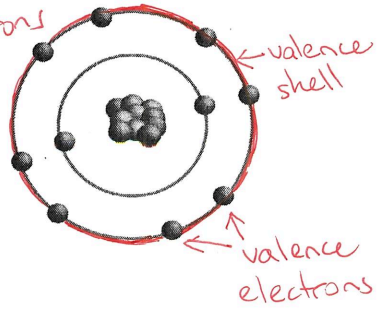
Bohr Models of Atoms and Ions

1. Calculate the number of protons, neutrons, electrons.
2. In the nucleus:
   - element symbol
   - # protons, # neutrons
3. Draw the electrons in energy shells:
   - Max electrons per shell from inside to outside: 2, 8, 8, (18)
   - (Except in first shell), electrons are filled ***starting at top***, going ***clockwise***, singly at first then paired
4. Ions only:
   - Add square brackets and ion charge from periodic table

Example: sodium atom	Example: oxygen ion
	

Valence Shell

- The **valence shell** is the outermost shell containing electrons
  Electrons in this shell are called valence electrons.
- A stable atom has a full valence shell.
- Atoms form compounds to have a full valence shell.
  - Ionic compound: atoms gain or lose electrons
  - Covalent compound: atoms share electrons.



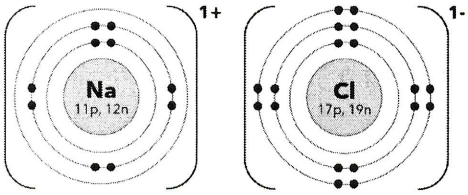
Ionic Compound Formation

- Atoms form ions to have a **full valence shell**, just like the noble gases have.
- Electrons are negatively charged. When electrons are added or taken away, atoms become positively or negatively charged ions.
  - **Cation**: positively charged ion (e.g. Ca^{2+}, Cr^{3+}, NH_4^+); forms when electrons are lost from an atom
  - **Anion**: negatively charged ion (e.g. N^{3-}, S^{2-}, PO_4^{3-}); forms when electrons are gained by an atom
- Ionic compounds form when **electrons are transferred** and ions are formed. Usually involves a **metal and a non-metal**.

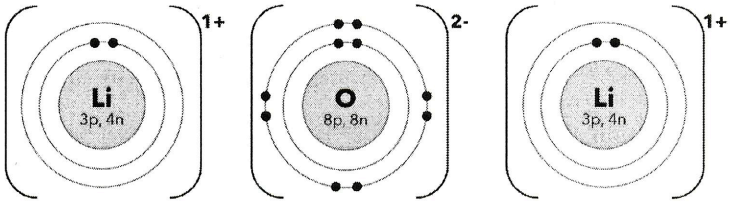
Bohr Models of Ionic Compounds

1. Determine how many of each ion is in the compound, from the subscripts.
2. Use the periodic table to find the ionic charge of each ion.
3. Draw the Bohr models of all the ions in the compound, side by side. (They should all have full valence shells.)

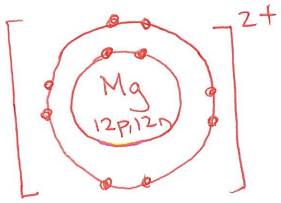
NaCl:

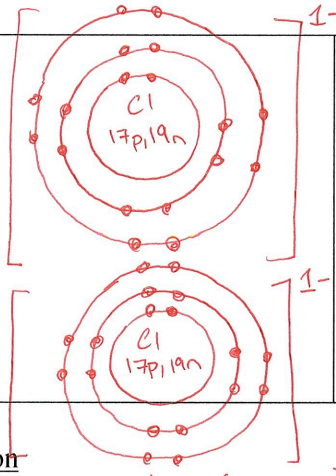


Li_2O:

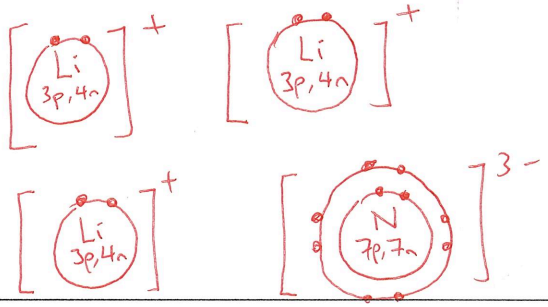


a) $MgCl_2$



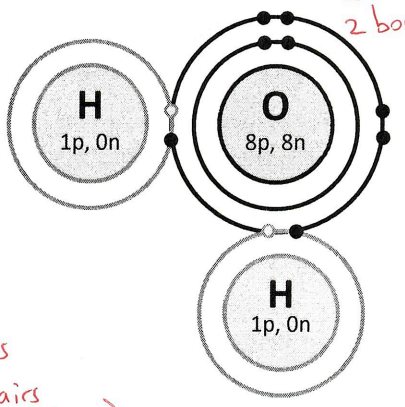


b) Li_3N

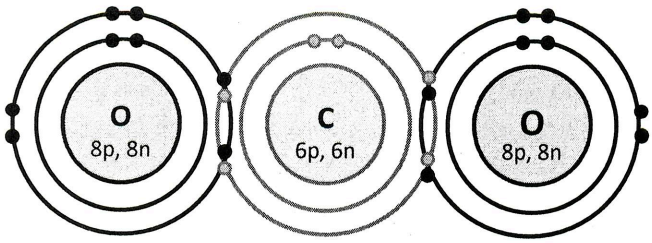


Covalent Compound Formation

- Covalent compounds form when two (or more) non-metal atoms share electrons.
- **Lone pair**: pair of valence electrons that is not shared between atoms
- **Bonding pair**: shared pair of valence electrons in a covalent compound



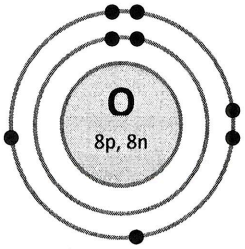
2 lone pairs
2 bonding pairs

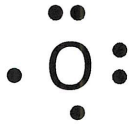


4 lone pairs
4 bonding pairs
(2 double bonds)

Introducing Lewis Structures

Bohr Model	Lewis Structure
• All electrons • All energy shells • Shows protons and neutrons • Shows a lot of information, but is clunky and time-consuming	• Only valence electrons (except cations) • Outermost shell only • Protons and neutrons ignored • Good at determining bonding in a covalent compound





Lewis Structures of Atoms

1. Write element symbol (capitalization matters!)
2. Draw valence electrons around, using the same positions as the Bohr model (i.e. clockwise, unpaired at first then paired)

Practice: *Draw the Lewis structures of:*

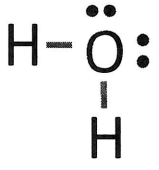
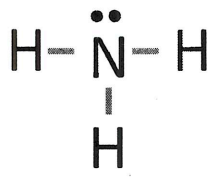
a) Mg atom	b) N atom	c) H atom	d) F atom
			

How do you figure out the number of valence electrons in an atom?

last digit of group #

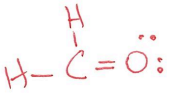
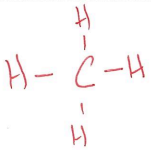
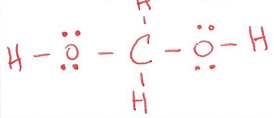
Lewis Structures of Covalent Compounds

Rule 1: All valence electrons (from the bonded atoms) must be used.
Rule 2: All atoms must have a full valence shell.

Example: H_2O	Example: NH_3
	

Steps:

1. Draw the Lewis structure of each atom. (Count how many valence electrons you have in total; write this down.)
2. Determine how many bonds each atom "needs" to complete its valence shell.
3. Guess and check with single, double, and triple bonds until your structure satisfies Rule 1 AND Rule 2.

HF	PF_3	CH_2O
		
N_2 *	CH_4	CO_2H_4 *(challenge)*
		

Diatomic Elements: When in their elemental form, exist as diatomic molecules: two atoms bonding covalently to fill their valence shells.

List: hydrogen, iodine, bromine, oxygen, nitrogen, chlorine, fluorine

$H_2, O_2, I_2, Br_2, N_2, Cl_2, F_2$

Memory Aid: HIBrONClF

Subscripts in Chemical Compounds

- Subscripts are small numbers written on the bottom right of an element or ion to show how many are in that compound.
- If there is no subscript, this means that there is only one of that element or ion.
- A subscript outside a bracket applies to the entire polyatomic ion inside the bracket (multiply subscripts!).

Circle all subscripts in the chemical equation below:

$$Cu + 2AgNO_{(3)} \rightarrow Cu(NO_{(3)})_{(2)} + 2Ag$$

Practice:

Chemical Formula	How many atoms?		Chemical Formula	How many atoms?		
Co_2S_3	Co:2	S:3	H_2O	H: 2	O: 1	
PF_4	P: 1	F: 4	CCl_4	C: 1	Cl: 4	
$MgBr_2$	Mg: 1	Br: 2	$CaCO_3$	Ca: 1	C: 1	O: 3
Be_3N_2	Be: 3	N: 2	NaOH	Na: 1	O: 1	H: 1

Practice:

Chemical Formula	Cation	Anion	Atom Count			
NaOH	Na^+ x1	OH^- x1	Na:1	O:1	H:1	
$Mg(OH)_2$	Mg^{2+} x1	OH^- x2	Mg:1	O:2	H:1	
$Be_3(PO_4)_2$	Be^{2+} x3	PO_4^{3-} x2	Be: 3	P: 2	O: 8	
$Ti_2(CrO_4)_3$	Ti^{3+} x2	CrO_4^{2-} x3	Ti: 2	Cr: 3	O: 12	
$(NH_4)_2Cr_2O_7$	NH_4^+ x2	$Cr_2O_7^{2-}$ x1	N: 2	H: 8	Cr: 2	O: 7

Reference

Non-metal Element	"-ide" Ending	Non-metal Element	"-ide" Ending
N, nitrogen	Nitride	**Se**, selenium	Selenide
O, oxygen	Oxide	**Br**, bromine	Bromide
F, fluorine	Fluoride	**I**, iodine	Iodide
P, phosphorus	Phosphide	**As**, arsenic *	Arsenide
S, sulfur	Sulfide	**Te**, tellurium *	Telluride
Cl, chlorine	Chloride	**At**, astatine *	Astatide

Arabic Numeral	Roman Numeral	Prefix
1	I	mono
2	II	di
3	III	tri
4	IV	tetra
5	V	penta
6	VI	hexa
7	VII	hepta
8	VIII	octa
9	IX	nona
10	X	deca