

EXPERIMENT 3

THE PERIODIC TABLE

EQUIPMENT:

CHEMICALS: None

Your lecture chemistry text/**review Lewis Dot Structures before lab**

Periodic Table

Graph paper

Calculators

Rulers

PURPOSE

In this experiment you will become familiar with some of the aspects of the periodic table of the elements. It can give you valuable information about the structures and properties of the elements as atoms, ions, and molecules. Rules for naming elements, compounds, and ions will be used as a means of gaining experience with the symbols that represent the composition of chemical substances. Lewis Dot Structures will be introduced as an exercise to examine bonding and structure. All of which is founded on the information in the Periodic Table.

DISCUSSION

The Periodic Table is an arrangement of the elements by electron configuration and increasing atomic number. They are arranged so that elements having the same number of outer electrons become the members of a family. These vertical columns are the “families of the Periodic Table.” This is useful because the number of outer electrons influences most of the chemical properties. These properties are said to be periodic because they change in a cyclic fashion with electron configuration. (The word periodic means recurring or cyclic). Examples of periodic properties are melting point, boiling point, atomic radius, oxidation number, ionization energy, and electronegativity. If one knows the periodic trends of these properties, one can determine relative values of these properties for various elements from their positions on the periodic table.

Other useful information about elements can be found easily from the periodic table. The name and/or symbol for each element, its atomic number, and its atomic mass are arranged in an easy to read manner. Frequently a Periodic Table will include the observed electronic configuration for each element.

The atomic number is the number of protons in an atom. For neutral atoms, this is equal to the number of electrons. Because the electron configuration of an atom is determined by the number of electrons, any property that is periodic with the electron configuration will be periodic with the atomic number. One part of this experiment is to graph a periodic property versus atomic number.

LEWIS DOT STRUCTURES

Abridged Lewis Dot Rules (see Appendix 1 in this lab manual for more)

1. Sum the valence electrons for all atoms in the molecule you want to draw (find each atom's periodic table group number and add together the group number of each atom in the compound). Next, add one electron to this number for each anionic charge; and subtract one for each cationic charge. The final number from this process is the number of valence electrons for the molecule you're drawing. Write this valence electron number down near the formula; you will need it later. Once you've established the number of valence electrons you can't change that number as you draw your Lewis dot structure.
2. For triatomic molecules, write down the atomic symbols for the atoms in the order they appear in the formula unless told to do otherwise. For molecules with four atoms or more, write the heteroatom in the center and bond all the other atoms to it. Join all atoms with one bond (2 electrons represented by a line joining atoms) to begin your structure.
3. Give all atoms in a structure an octet using additional bonds (lines) or lone pairs of electrons (dots). Hydrogen always only shares two electrons. Don't add more electrons than the number you wrote down in **Step 1**.
4. Since every atom in the structure gets an octet (except hydrogen), use multiple bonds between atoms to avoid adding more electrons than you wrote down in **Step 1** to make sure all atoms get an octet. When you think you're finished, count all your bonds (2 electrons) and lone pairs (dots) and add them up. Write down this number; it cannot exceed the number you wrote down in **Step 1**. If it does....you made a mistake. Erase all your lone pairs and go back to **Step 3** and start again.

Each element is found in a group in the periodic table: chlorine (Cl) is in group VII (17) and aluminum (Al) is in group III (13) of the representative elements. This group number (or group number -10) indicates the number of valence electrons or reactive frontal electrons. Chlorine has seven valence electrons and aluminum three. As we have learned in lecture, the metallic elements are electropositive, tending to lose their valence electrons to form cations, and the nonmetals are electronegative, tending to attract enough electrons to make a total of eight electrons.

To “draw” a Lewis Dot structure for an element, one must indicate the symbol for the element, and then put some dots, equal to the group number, around the symbol forming a square if the element has four or more electrons. Examples:

Group V (15)
valance electrons = $5e^-$



Group VI (16)
valance electrons = $6e^-$



Group VII (17)
valance electrons = $7e^-$

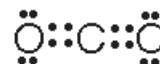



Lewis dot structures are most useful in determining the bonding in molecular compounds like HCN as opposed to ionic compounds, like CaBr_2 . In molecular compounds, electrons are “shared” by both nonmetals so that the atoms can each obtain the equivalent of eight valence electrons (an octet). The “shared” electron pairs are bonding electrons and are counted twice, once in the grouping for each nonmetal atom in the bonding interaction. The other electrons are “non bonding” electrons and are located in pairs directed away from the bonds. At times, one must “share” more than one pair of electrons to make the required 8 & 8. This represents the multiple bonding observed in diatomic, molecular oxygen and nitrogen. Examples are shown below:

for bromine, Br_2
valence electrons = 7 + 7 = 14e⁻



for carbon dioxide, CO_2
valence electrons = 4 + 6 + 6 = 16e⁻



for nitrate, NO_3^{1-} 
valence electrons = 5 + 3(6) + 1 = 24e⁻



Note nitrate's “additional” electron from the -1 charge is shown with brackets when drawing the Lewis dot structure. All ions have the brackets and charge displayed.

NOMENCLATURE

In addition to periodic properties of elements, we have provided some rules and examples for naming chemical compounds, elements, and ions. These rules are summarized in a flow chart, Table II. This flow chart is not intended to replace the memorization of nomenclature rules, but it relates the names of the compounds to the chemical properties of the elements as indicated by their location in the periodic table. It should be used as an aid along with those rules. With practice you should be able to name substances from their chemical formulas or write correct formulas from their names.

$\text{Br}_2\text{I}_2\text{N}_2\text{Cl}_2\text{H}_2\text{O}_2\text{F}_2$ What!!!!

An easy way to remember which of the elements in nature occur in **diatomic** forms is to memorize the “word” BrINClHOF. (Sounds kinds like a German baker or peanut brittle, eh?) All of the chemical symbols in BrINClHOF—bromine, iodine, nitrogen, chlorine, hydrogen, oxygen and fluorine—exist **elementally** as naturally occurring diatomic molecules with oxidation state zero. Most other elements (sulfur, S_8 , is an exception) are **monatomic** in their elemental form in nature. This means when some one says “oxygen in the atmosphere” that person means elemental oxygen, O_2 . When one says “metallic iron” it means elemental iron, **Fe**. Both of these are the elemental forms with oxidation state zero.

TABLE I**SELECTED IONS AND THEIR NAMES****Cations with one oxidation state**

H^+	hydrogen	Mg^{2+}	magnesium
Li^+	lithium	Ca^{2+}	calcium
Na^+	sodium	Sr^{2+}	strontium
K^+	potassium	Ba^{2+}	barium
Ag^+	silver	Zn^{2+}	zinc
NH_4^+	ammonium	Al^{3+}	aluminum

Cations with more than one oxidation state

Cr^{2+}	chromium(II)	Cr^{3+}	chromium(III)
Cu^+	copper(I)	Cu^{2+}	copper(II)
Fe^{2+}	iron(II)	Fe^{3+}	iron(III)
Pb^{2+}	lead(II)	Pb^{4+}	lead(IV)
Mn^{2+}	manganese(II)	Mn^{3+}	manganese(III)
Hg_2^{2+}	mercury(I) (dimer)	Hg^{2+}	mercury(II)
Ni^{2+}	nickel(II)	Ni^{3+}	nickel(III)
Sn^{2+}	tin(II)	Sn^{4+}	tin(IV)

Notice no space

TABLE II**SIMPLE ANIONS**

H ⁻	hydride	O ²⁻	oxide
F ⁻	fluoride	S ²⁻	sulfide
Cl ⁻	chloride	O ¹⁻	peroxide (or O ₂ ²⁻)
Br ⁻	bromide	N ³⁻	nitride
I ⁻	iodide	P ³⁻	phosphide
C ₂ H ₃ O ₂ ⁻	acetate	HSO ₄ ⁻	hydrogen sulfate or bisulfate
HCO ₃ ⁻	hydrogen carbonate or bicarbonate	HSO ₃ ⁻	hydrogen sulfite or bisulfite
CrO ₄ ²⁻	chromate	CO ₃ ²⁻	carbonate
CN ⁻	cyanide	ClO ₃ ⁻	chlorate
Cr ₂ O ₇ ²⁻	dichromate	ClO ₂ ⁻	chlorite
OH ⁻	hydroxide	PO ₄ ³⁻	phosphate
ClO ⁻	hypochlorite	PO ₃ ³⁻	phosphite
NO ₃ ⁻	nitrate	H ₂ PO ₄ ⁻	dihydrogen phosphate
NO ₂ ⁻	nitrite	HPO ₄ ²⁻	hydrogen phosphate
C ₂ O ₄ ²⁻	oxalate	SO ₄ ²⁻	sulfate
ClO ₄ ⁻	perchlorate	SO ₃ ²⁻	sulfite
MnO ₄ ⁻	permanganate	SCN ⁻	thiocyanate
S ₂ O ₃ ²⁻	thiosulfate		

FLOW CHART FOR NAMING COMPOUNDS

1. Does the compound contain metallic elements?

If yes, go to 2.

If no, go to 3.

2. Does the metal have a single oxidation state?

If yes, go to 4.

If no, go to 5.

3. Is the nonmetallic compound an acid? Does it contain H and other nonmetals?

If yes, go to 6.

If no, go to 7.

4. An ionic compound: The compound's name is formed from the metal cation name and the anion name. **The cation name goes first!**

example: sodium sulfate — Na^+ and SO_4^{2-} as Na_2SO_4

aluminum hydroxide — Al^{3+} and OH^- as $\text{Al}(\text{OH})_3$

Note: You need two sodiums and three hydroxides to compensate for the charges on the counter ion.

5. An ionic compound: The cation's oxidation state is given as a Roman numeral after the metal cation's name. This is followed by the anion name. Again, ionic compound's name just includes the ion's names.

example: iron(III) chloride — Fe^{3+} and Cl^- as FeCl_3

6. An acid: Does the anion in this hydrogen-anion compound contain oxygen? Is it an "oxy-acid" or a "binary acid"?

If yes, the ion's suffixes -ate and -ite on the anion are replaced by -ic and -ous, respectively, to form the oxyacid's name.

examples: sulfurous acid — H_2SO_3 from SO_3^{2-} , sulfite

If no, the name for the binary acid includes the prefix hydro- and suffix -ic and is derived from the monatomic anion's name (which ends in -ide).

examples: sulfuric acid — H_2SO_4 from SO_4^{2-} , sulfate

hydrochloric acid — HCl from Cl^- , chloride

7. Binary, nonmetallic compound: The most electropositive element goes first (see page 34 or your lecture text). The most electronegative element gets the suffix -ide. Use prefixes to indicate the number of each element. These are the same ones we know from the “angles” and “gons,” like triangles and pentagons (polygons).

example: diphosphorous pentoxide — P_2O_5

8. If the first element in the formula (the least electronegative—most electropositive) is a single atom then the prefix mono is left off.

example: carbon monoxide — CO
(not monocarbon monoxide)

$Br_2I_2N_2Cl_2H_2O_2F_2$ What!!!!

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Your Name _____ TA's Name _____ Lab Section 1411. _____ Date _____

PROCEDURE

10%

PART A. PERIODIC PROPERTIES

Make **one plot** that includes data of one of the periodic properties versus atomic number—see data at end of this lab (don't use Group number). Use the atomic number as the x-axis and the value for the property as the y-axis. To do this, you must first calculate how many units of that property are represented by each square of the graph paper (blank graph papers are found at the end of this lab manual). Label the y-axis accordingly. Label the x-axis with the chemical symbols for the atomic numbers 1-38 (example N = 7; Ni = 28). Plot each point for your property by placing a small circle for each atomic number at the corresponding value of the property. Connect these points with a line. Note the cyclic nature of your graph.

10%

STUDY QUESTIONS

1. Compare the portion of the graph that correspond to the first period (lithium through neon) with the portion that corresponds to the second period (sodium through argon). Compare with the portion for the third period and the fourth period. How is your graph cyclic?
2. Which family (group not period, for instance "which oxygen family" or "which Group 2 family") has the highest value for the property you plotted? Which has the lowest?
3. Does the property plotted increase or decrease with atomic number?

Graph paper is in the back of this lab manual.

PART B: CHEMICAL COMPOUNDS

Fill in the correct formula and name for each compound.

Use the ionic nomenclature described on pages 27 and 28.

Ionic Compounds

	Cl ¹⁻ <i>example</i>	NO ₃ ¹⁻	SO ₄ ²⁻	PO ₃ ³⁻
Na ⁺	NaCl sodium chloride			
NH ₄ ⁺				
Al ³⁺				
Cu ²⁺				
Pb ⁴⁺				
Mn ³⁺				

20%

Covalent Compounds

Name the following covalent compounds. (These are not ions.):

SO₂ _____

↑
important

SO₃ _____

NO _____

NO₂ _____

P₂S₅ _____

20%

Problems similar to this will be on the scantron-based midterm and final exams. If you can successfully complete these you should do well on those tests.

Look at the naming procedures on pages 29 and 30 for help.

Your Name _____ TA's Name _____ Lab Section 1411. ____ Date _____

PART C: LEWIS DOT STRUCTURES

Use the Lewis Dot structure rules on page 92 of this lab manual for the following:

Write Lewis electron dot structures for each of these diatomic gases:

valence electron # = $2(1) = 2$



valence electron # = $2(5) = 10$



valence electron # =



valence electron # =



Covalent compounds (**these are not charged chemical species**):

valence electron # =



valence electron # =



valence electron # =



valence electron # =



valence electron # =



valence electron # =



valence electron # =



Only
expanded
octet on
page

Ions (these HAVE charges; neither has an expanded octet):

valence electron # =



valence electron # =



Problems similar to this will be on the scantron-based midterm and final exams.
If you can successfully complete these you should do well on those tests.

PERIODIC PROPERTIES OF THE ELEMENTS

Symbol	Atomic Number	Group Number	Ionization Potential (eV)	Electron Affinity (kJ/mol)	Atomic Radius (1x10 ⁻¹⁰ m)	Electro-negativity
H	1	1	13.6	0.75	0.37	2.2
He	2	8	24.6		0.49	
Li	3	1	5.4	59.8	1.52	
Be	4	2	9.3		1.11	1.6
B	5	3	8.3	26.8	0.88	1.8
C	6	4	11.3	122.3	0.77	2.5
N	7	5	14.5		0.7	3
O	8	6	13.6	141	0.66	3.4
F	9	7	17.4	328	0.64	4
Ne	10	8	21.6		0.7	
Na	11	1	5.1	52.7	1.86	0.93
Mg	12	2	7.6		1.6	1.3
Al	13	3	6	42.6	1.43	1.6
Si	14	4	8.1	133.6	1.17	1.9
P	15	5	10.5	71.7	1.1	2.2
S	16	6	10.4	200.4	1.04	2.6
Cl	17	7	13	348.8	0.99	3.2
Ar	18	8	15.8		0.94	
K	19	1	4.3	48.4	2.31	0.82
Ca	20	2	6.1		1.97	1
Sc	21		6.5			1.4
Ti	22		6.8	7.7		1.5
V	23		6.7	50.8		1.6
Cr	24		6.8	64.4		1.7
Mn	25		7.4			1.6
Fe	26		7.8	24		1.8
Co	27		7.9	67		1.9
Ni	28		7.6	111		1.9
Cu	29		7.7	1111		2
Zn	30		9.4			1.6
Ga	31	3	6	29	1.22	1.8
Ge	32	4	7.9	116	1.22	2
As	33	5	9.8	77	1.21	2.2
Se	34	6	9.8	195	1.17	2.6
Br	35	7	11.8	324.6	1.14	3
Kr	36	8	14		1.09	
Rb	37	1	4.2	46.9	2.44	0.82
Sr	38	2	5.7		2.15	0.9