Periodic Table – Unit ________ Name______________________________ 1

I. Introduction to the Periodic Table –

A. Mendeleev's Periodic Table – Dmitri Mendeleev in 1869 developed a novel system for classification of elements.

- He organized the elements by increasing atomic mass.

B. Modern Periodic Table – Henry Moseley in 1913 determined an atomic number for each known element

- This information served as the basis for the modern periodic table
- The modern periodic table is organized by elements with increasing atomic number!

C. Arrangement of Periodic Table

GROUPS (columns) OR FAMILIES

- Elements within a *group* have the same # of valence electrons
- Elements within a *group* have similar properties. (react similarly)

PERIODS (rows)

Elements within a *period* have the same # of PELs OR shells

Elements within a *period* have different properties.

Ex: Name an element that would have chemical properties similar to Mg: Ca (same group up + down)

D. Periodic Table Basics – where are the elements located? The group names and locations that you must memorize are:

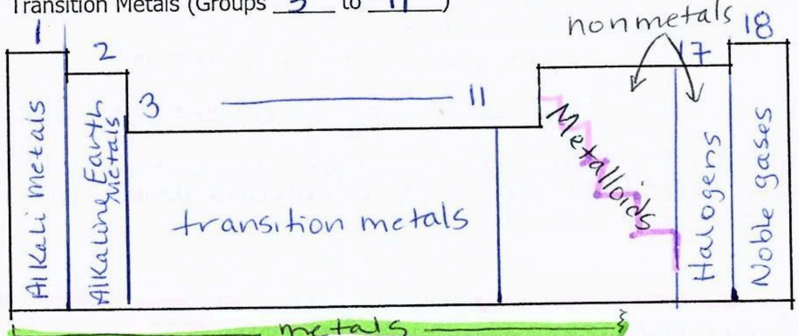
Alkali Metals (Group 1)

Alkaline Earth Metals (Group 2)

Transition Metals (Groups 3 to 11)

Halogens (Group 17)

Noble Gases (Group 18)



1. Metals – Metals are "losers" of electrons – only a few valence electrons loosely held

- physical properties:

malleable (hammered into sheets) metallic luster (shiny)

ductile (made into wire) conduct heat & electricity

- most metallic element: Fr Only liquid metal: Hg (Mercury)

2. non-metals – Think of wax — includes Hydrogen

Carbon is a nonmetal

- physical properties:

dull, non-conductors, brittle

- most non-metallic element: F Only liquid non-metal: Br

3. metalloids – or Semi-metals
- physical properties:

properties of both metals + nonmetals

- list the 7 metalloids: B, Si, Ge, As, Sb, Te, At

(on stair-case)

Aluminum + Polonium are "metallic"

II. Information given on NYS Reference Tables for Physical Setting/CHEMISTRY

A. Periodic Table – p. 9: Copy all information in the box for Carbon:

average atomic mass → 12.0111 (NOT mass #)

12.0111	−4 +2 +4
C	
6 2-4	

oxidation states – or charges of atom when it bonds

atomic # → 6

↑ Electron configuration (Bohr)

Define or explain the following:

12.0111 u Atomic mass – weighted average of all possible isotopes for that element

$_6C$ Atomic Number – total number of protons

2 − 4 Electron Configuration – number of electrons per shell

C^{+4} Oxidation State – possible charges an atom will get when losing or gaining electrons when bonding

B. Table S – Properties of Selected Elements

(IE)

1. Atomic number, symbol, name – are listed in order of atomic number
2. First Ionization Energy – define and **memorize:** the number of kJ/mol Kilojoules per mole of energy required to remove the outermost electron

Table S

metals want to lose electrons

- metals: have low ionization energy values
- which means they tend to lose electrons easily
- forming positive charged ions

Na• → Na

2-8-1 2-8

it requires 496 KJ to remove the 1 e⁻ from 1 mole of Sodium atoms

:F:

F does not want to give away an e⁻ so it takes more energy to remove it. IE = 1681 KJ/mol (nonmetals have high IE values)

2

Fluorine atom :F: + 1e⁻ → [:F:]⁻ Fluoride ion

- non-metals: have high ionization energy values
- which means that non-metals tend to gain electrons rather than lose them, and that they form negative ions.

(EN) 3. Electronegativity – define and memorize: the ability or desire of an atom to gain an e⁻
only a scale (0–4); not a measurement

- metals: have low (low or high) electronegativity values; this indicates that they have a weak (strong or weak) attraction for bonded electrons.

Think of F the no. (4.0)

- non-metals: have high (low or high) electronegativity values; this indicates that they have a strong (strong or weak) attraction for bonded electrons.

4. Melting Point, Boiling Point – in units of Kelvin. ~~Is the melting point an intensive or an extensive property? Explain.~~
melting and freezing points are identical

melting
0°C
freezing
H_2O

5. Density – define: how tightly packed the atoms are $d = \frac{mass}{volume}$
$(1\,cm^3 = 1\,mL)$

(AR) 6. Atomic Radius – define half the distance between 2 nuclei
H AR = 32pm (32 picometers)
32×10^{-12} meter

III. Group Information – found in chapter 6 and handout

Groups are organized in columns on the periodic table. The group number is associated with the number of valence electrons.

A. Representative elements:

- Representative elements include all elements in groups 1, 2, 13–18
- Atoms of representative elements are filling the s and p sublevels
- Where is the "s" block? groups 1 + 2
- Where is the "p" block? groups 13–18

Group 1

1. Name of this group? alkali metals
2. Which element in this group is *not* a metal? H Why is it included in this group? atomic # of 1, forms +1 ions, it's also the tiniest element

$Na \cdot \rightarrow Na^{+1}$
2-8-1 → 2-8

3. Describe the reactivity of these elements. highly reactive

4. How are these elements found in nature? only bonded in nature

NaCl NaF LiBr

Why? only need to lose 1e⁻ to have a full shell of 8e⁻

5. How many valence electrons does this group have? 1

6. Describe the ionization energies and electronegativities of this group (are they low or high?) Table S Low, low

What does this mean? Low IE - little energy required to remove an e⁻. Low EN - low desire to gain an e⁻

Group 2

1. Name of this group? alkaline earth metals

2. Describe the reactivity of these elements. highly reactive but not quite as high as group 1

3. How are these elements found in nature? only bonded

Why? only need to lose 2e⁻ to be stable

$CaCl_2$ MgF_2

4. How many valence electrons does this group have? 2

Group 13

1. How many valence electrons does this group have? 3

2. This first element in this group, B, is a semimetal. All of the other elements in this group are metals.

Group 14

1. How many valence electrons does this group have? 4

2. This group has C, which is a non metal, Si and Ge are semimetals and Sn, Pb which are metals.

3. What is an allotrope? Define: same element but different forms so they have different properties

4. What are the allotropic forms of carbon? How are they different from each other?

Pure carbon can be found as a diamond or as graphite
diamond is hard, graphite is soft

Oxygen:
O_2 we breathe this
:O=O:

O_3 ozone - toxic gas

Group 15

1. How many valence electrons does this group have? 5
2. This group also makes a transition from non-metal to semimetal to metal.
3. What is interesting about nitrogen?  :N≡N: quite stable due to triple bond
   - It makes up 76% of the atmosphere
   - Importance of nitrogen and phosphorous in living things:

N_2 is basic ingredient in amino acids
P turns nutrients into energy

Group 16

1. How many valence electrons does this group have? 6
2. Describe the reactivity of oxygen: reactive (think of metal in "air")
3. Why is there so much oxygen in the atmosphere? produced by plants during photosynthesis
4. Name two allotropes of oxygen: oxygen and ozone. Do allotropes have similar or different properties? different? Why? different structures

O_2 oxygen :O=O:

O_3 ozone 

Group 17

1. What is the name of this group? halogens
2. How many valence electrons does this group have? 7
3. Describe the reactivity of these elements: very reactive – only found combined in nature Cl_2 NaCl
4. What kind of electronegativities and ionization energies does this group have? very high
   What does this mean? high desire to gain an e- (high EN) high IE – it takes a lot of energy to remove an e-
5. Which group 17 element has the highest electronegativity? F
6. Describe the change in phase (solid, liquid or gas) takes place going down the column from fluorine to iodine:

Fluorine is a gas
Bromine is a liquid
Chlorine is a gas **
Iodine is a solid

I_2 is much bigger than F_2 or Cl_2 so it has stronger forces between the molecules

Group 18 Rn (Radon) "silent Killer" Radon gas

1. What are elements in this group called? noble gases

2. How many valence electrons does this group have? 8 (except Helium which has 2)

3. Each of these elements exist as monatomic gases.
   Monatomic: 1 atom Ne, Ar, Xe . . .

(compared to diatomic gases H_2, F_2, Cl_2 . . .)

4. What is special about this group's ionization energy values?
   Very high - does not want to give away e− you must supply a lot of energy to remove e−

5. Why is the term "inert" no longer applicable?
   F and O can react with the big, heavy, lazy noble gases

B. Transition Elements – group numbers ________ to ________; between "s" and "p" block

1. Within the table
   - These elements are characterized by having electrons in ______ orbitals
   - Therefore this block is called the "_______" block
   - Ions of transition metals produce colored solutions (see pictures on pages 168 – 169 in your text)
   - Several transition metals have more than one possible oxidation states.
   - Examples of transition metals:
     $CuSO_4(aq)$ is a pretty blue solution

2. The bottom two rows on your periodic table (the cut out rows)
   - These are called ________________ transition elements
   - These elements are filling the _____________ sublevel.
   - Therefore this little block is known as the " ______" block.

Give two interesting facts about the lanthanide series:

Give two interesting facts about the actinide series:

This is the end of the HW assignment! Keep going ⟹

IV. Periodic Table Trends Use Table S

A. Group/Family Trends: Trends as you go down a group ↓

1. Atomic radius: going down a group the atomic radius increases
Why? due to the addition of a new shell or energy level with each row

2. Electronegativity: going down a group the electronegativity decreases
Why? less desire for an e-, valence shell is further from nucleus, less nuclear pull or attraction for e-. Inner electrons also shield valence shell from nuclear pull.

3. Ionization energy: going down a group the ionization energy decreases
Why? less nuclear pull - it takes less energy to remove e- from valence shell because the valence electrons are further from the nucleus

SUMMARY: Each of these trends (1-3) can be explained in terms of:
adding a shell going down the group

4. Metallic/nonmetallic properties (what is happening to the properties of the elements as you go down a group?)
↓ bottom is more "metallic"
most "metallic" element is Fr (loses e- easiest)

B. Periodic Trends: Trends as you go across a period from L→ R

1. Atomic radius: going across a period the atomic radius decreases
Why? electrons going into same shell - increased number of protons so nuclear pull is greater as you go left to right (see Table S)

2. Electronegativity: going across a period the EN increases
Why? nonmetals have more valence electrons - closer to 8 so their "desire to gain" another electron increases.

3. Ionization energy: going across a period the ionization energy increases
Why: closer to 8 valence electrons - atom does not want to lose any electrons so more energy must be used to remove an e-

SUMMARY: Each of these trends (1-3) can be explained in terms of:
nuclear pull increases Left to Right

4. Metallic/nonmetallic properties (what is happening to the properties of the elements as you go across a period?)

→ more nonmetallic
most "nonmetallic" element is F (gains e- easiest)

V. Other

A. Comparing the size of an atom to an ion:

Metals:

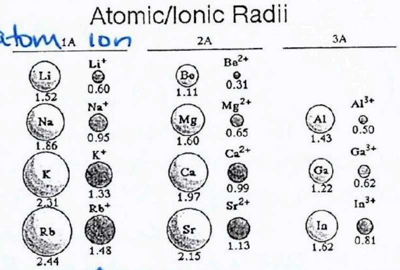
Metals lose electrons when bonding

Na	Na^+
2-8-1	2-8
atom	ion (only 2 shells)

thus Metallic ions are smaller than their atoms

Atomic/Ionic Radii

atom ion

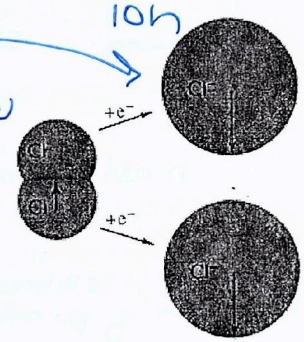


Nonmetals are the opposite

nonmetal ions are larger than their atoms

ion

atom



(atom) Covalent radius 99 pm

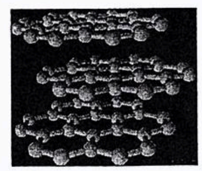
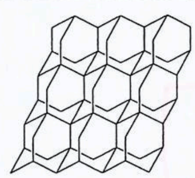
Anionic radius 181 pm (ion)

B. Isoelectronic: same electronic structure / similar

Mg^{2+}	Na^+	Ne	F^-	O^{2-}	
2-8	2-8	2-8	2-8	2-8	← e⁻ configuration for the ions.
12p+	11p+	10p+	9p+	8p+	
10e⁻	10e⁻	10e⁻	10e⁻	10e⁻	
2-8-2	2-8-1	2-8	2-7	2-6	← Bohr e⁻ configuration for the atoms

Extension Questions

Model 4 – Allotropes of Carbon

Natural Sample	Properties	Structure	Composition
 Graphite	Black Soft Conductive		98.89% Carbon-12 1.11% Carbon-13 pure Carbon
 Diamond	Colorless Very hard Insulator		98.89% Carbon-12 1.11% Carbon-13 pure Carbon

18. Consider the information about carbon provided in Model 4.

    a. Are diamonds and graphite made from the same element? yes

    b. Can the existence of isotopes explain the difference in properties between diamond and graphite? Explain. no - notice the percentages above

    c. Propose an explanation for the difference in properties between diamond and graphite.

    due to the arrangement of the Carbon atoms

19. O_2 and O_3 (ozone) are allotropes of oxygen. Buckminsterfullerene (C_{60}) is another allotrope of carbon. Based on these statements and the information in Model 4, propose a definition for **allotrope**. forms of the same element with different atom arrangement, thus different properties

20. Two common forms of phosphorus are red and white. Red phosphorus is fairly stable at room temperature in air, but white phosphorus can ignite easily when exposed to air. Is this difference in properties due to the existence of different isotopes of phosphorus or different allotropes? Explain.

    "Phosphorus" contains many different isotopes ^{28}P ^{29}P ^{30}P ^{31}P ^{32}P