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Study Guide: Chemistry Semester 1

ATOMIC & MOLECULAR STRUCTURE STANDARD 1

Atomic number

1. What is **atomic number**?

The # of protons in the nucleus

2. What is atomic number for the following elements?

Cr	Iron <i>26</i>
Carbon <i>6</i>	Ca <i>20</i>

3. How does atomic number change moving left to right across the periodic table? *it increases*

Atomic Weight (Atomic Mass)

4. What is the **atomic mass** for the following elements?

V <i>50.94g/mol</i>	Oxygen <i>16.00g/mol</i>
Beryllium <i>9.01g/mol</i>	Mn <i>54.94g/mol</i>

5. It was found that the periodic table was organized by increasing atomic number rather than mass. Give an example of a pair of elements where one element has a larger atomic number but a smaller atomic mass.
 $^{18}\text{Ar} \rightarrow 39.95 \text{ g/mol}$ $^{39}\text{K} \rightarrow 39.10 \text{ g/mol}$

6. Identify the letters in the following periodic square:

A = Atomic #
B = Atomic mass
C = Symbol
D = Name

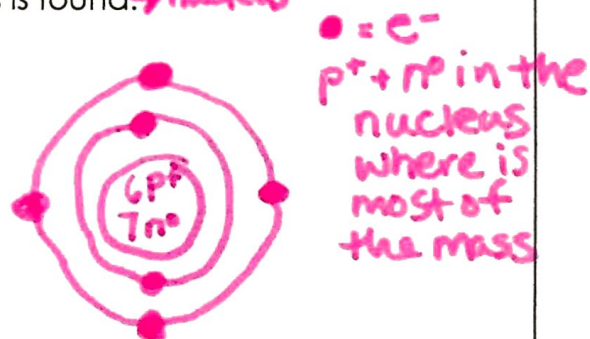
A →	16	
	S	← C
B →	32.066	
	Sulfur	← D

Atomic structure

1. Draw an atomic model of **Carbon-13**.

In the drawing include the following:

- Location of nucleus & electron cloud.
- Location of 3 subatomic particles.
- Amount of each subatomic particles.
- A note where most of the atom's mass is found. → *nucleus*

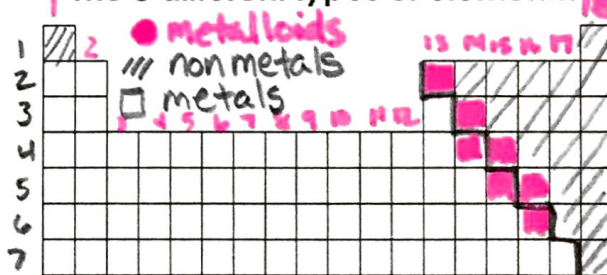


6p+ 7n0 6e-

Organization of the Periodic Table

1. Identify on the periodic table below the location of

The 3 different types of elements.



2. Where are **groups** on the periodic table? *# 1 → 18 vertical columns*

3. Where are **periods** on the periodic table? *# 1-7, horizontal rows*

Periodicity (periodic trends)

7. Atomic Radius

- Which element has the largest atomic radius?

Francium

- Which element has the smallest atomic radius?

Helium

8. Ionization energy

- What is Ionization Energy?

The energy required to remove an e^- from an atom

- Which element has the largest ionization energy?

Helium

- Which element has the lowest ionization energy?

Francium

- Why does ionization energy decrease moving down a group?

Each step down a group adds an extra energy level of electrons in the electron cloud shielding the outer electrons from the nucleus. Also, the larger the atom the lower the Coulombic Attraction.

9. Electronegativity

- What is Electronegativity?

An atom's ability to attract e^- s in a chemical compound

- Which element has the largest electronegativity?

Fluorine, F

- Which element has the lowest electronegativity?

Francium, Fr

- Why do Noble Gases tend to have no electronegativity?

They are stable and do not want to attract e^- s

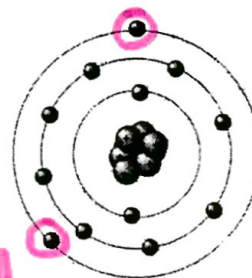
Valence electrons

10. Why are alkali metals so reactive?

They only have 1 valence e^- to lose to gain the Noble gas stability

11. Why are halogens so reactive?

They only need 1 valence e^- to fill their outer shell



12. Identify the amount of valence electrons in each of the following:

Na	1	Phosphorus	5
Argon	8	He	2

5. How many of magnesium atom's 12 electrons are valence electrons? Circle those electrons. 2

Review Chemistry I: Ionic Bonding and Nomenclature

1. Positive ions are called cations. Negative ions are called anions.

2. Why do chemical bonds form between atoms? (or...Why don't atoms just stay single?)

So each atom can fill their outer shell and become more stable

3. In an ionic bond, electrons are shared / transferred / connected between atoms. (circle one)

4. An ionic bond is a chemical bond caused by electrostatic attraction between cations and anions that is formed by transferring electrons between atoms.

5. Ionic bonds form between metals and nonmetals (metals, nonmetals, metalloids)

6. For the pairs of elements listed below, circle pairs that would likely form ionic bonds.

C and H Na and F Hg and Ag Mg and S N and C K and O

7. What is a polyatomic ion?

A group of atoms w/ a positive or negative charge

8. Circle the ionic compounds listed below.

CaSO₄ CH₄ CO₂ BaCl₂ NH₃ KNO₃ LiOH

9. How did you know which compounds in Question 8 above were ionic?

metal bonded w/ a nonmetal or metal bonded w/ a polyatomic ion

10. What are binary ionic compounds?

Ionic compounds consisting of only 2 elements

11. Binary ionic compounds typically end with the suffix -ide.

12. A "formula unit" is the simplest or lowest ratio of ions in an ionic compound.

13. How do each of the following atoms achieve a stable octet?

(circle gain or lose and write in a number of electrons)

O will gain / lose 2 electrons

Li will gain / lose 1 electrons

N will gain / lose 3 electrons

14. List the oxidation number (charge) for the following groups of elements.

Group 1	<u>+1</u>	Group 15	<u>-3</u>
Group 2	<u>+2</u>	Group 16	<u>-2</u>
Group 3	<u>+3</u>	Group 17	<u>-1</u>

15. List 3 physical properties of ionic compounds:

1. State: solid / liquid / gas (circle)

2. Melting Point: high / low (circle)

3. Conductor of electricity when melted or dissolved

16. Determine the total number of atoms in each formula listed below.

Li_3PO_4

8

$\text{Ba}(\text{NO}_3)_2$

9

17. Determine the total number of ions in each formula listed below.

Li_3PO_4

4

$\text{Ba}(\text{NO}_3)_2$

3

18. Which of the following compounds contains the Sn^{4+} ion?

A. Sn_2Br

B. SnCl

C. SnI_2

D. SnO_2

19. Name the following compounds.

MgI_2

magnesium iodide

Fe_2O_3

iron(III) oxide

$\text{Al}(\text{NO}_3)_3$

aluminum nitrate

NH_4Br

ammonium bromide

CuCO_3

copper(II) carbonate

KHCO_3

potassium bicarbonate

$\text{Ba}(\text{C}_2\text{H}_3\text{O}_2)_2$

barium acetate

SnCl_4

tin(IV) chloride

20. Write a formula for the following chemical compounds.

copper(I) oxide

Cu_2O

ammonium nitrate

NH_4NO_3

potassium phosphate

K_3PO_4

I

Iron(III) sulfide

Fe_2S_3

calcium hydroxide

$\text{Ca}(\text{OH})_2$

lead(II) iodide

PbI_2

aluminum sulfate

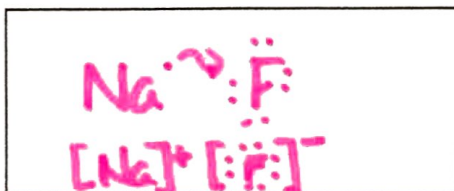
$\text{Al}_2(\text{SO}_4)_3$

nickel(II) bromide

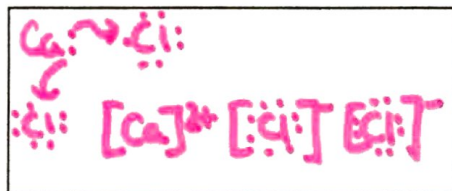
NiBr_2

21. Draw a Lewis Dot diagram for each the following ionic compounds.

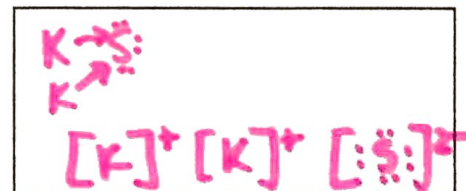
NaF



CaCl_2



K_2S



22. What is the oxidation number of the unknown element X in the compound MgX_2 ?

A. 1-

B. 2-

C. 1+

D. 2+

Review Chemistry I: Covalent Bonding and Nomenclature

1. Why do atoms share electrons in covalent bonds with other atoms rather than remaining as single atoms?

Atoms form covalent bonds to obtain a more stable electron arrangement → to obtain an octet

2. A molecule is a neutral group of atoms held together by _____.

- A. ionic bonds
B. unshared electrons
C. partial charges
D. covalent bonds

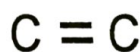
3. Compounds formed by covalent bonds usually contain _____.

- A. halogens and oxygen
B. two or more nonmetals
C. a metal and a nonmetal
D. positive and negative ions

4. Which of the following is a molecular compound?

- A. Na_2O
B. AlCl_3
C. SCl_6
D. CuO

5. How do atoms form covalent double bonds?



- A. One atom loses two electrons to the other atom in the bond.
B. Two atoms share two pairs of electrons.
C. Two atoms share two single electrons.
D. Two atoms share one electron.

6. Which of these elements does NOT exist as a diatomic molecule?

- A. P
B. Cl
C. O
D. N



7. A bond formed between a carbon atom and hydrogen atom is likely to be _____.

- A. single, covalent
B. double, covalent
C. ionic
D. triple, covalent

8. Which of the following bonds is the most polar (i.e. has the greatest difference in electronegativity)?

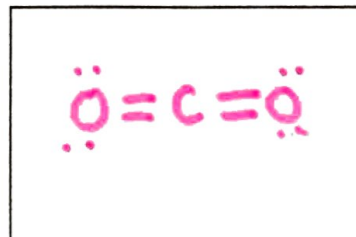
- A. C—C
B. H—N
C. O—H
D. H—Cl

9. Which of the following covalent bonds is the least polar?

- A. H—Cl
- B. H—C
- C. Cl—Cl**
- D. H—N

10. Which is TRUE of a nonpolar covalent bond?

- A. electrons are shared unequally between atoms
- B. a cation is bonded to an anion
- C. electrons are transferred between atoms
- D. electrons are shared equally between atoms**



11. Draw the Lewis dot structure for CO₂ in the box

12. Describe properties that are common in typical molecular compounds.

conductor: good or **poor**

melting point: high or **low**

boiling point: high or **low**

net charge: **neutral** or charged

#13-18 Write the NAME or FORMULA for the following molecular compounds:

13. CS₂ carbon disulfide

14. CCl₄ carbon tetrachloride

15. S₂O₆ disulfur hexoxide

16. triboron monoxide B₃O

17. tetraphosphorus pentoxide P₄O₁₀

18. arsenic trihydride AsH₃

19. In drawing a Lewis structure, what is special about hydrogen (H) ?

- A. It achieves an octet of 8 electrons.
- B. It can form up to four bonds with 8 shared electrons.
- C. It can only form one double bond.
- D. It can only form one single bond with no unshared electrons around it.**

For #20-25:

-Draw the Lewis dot structures for the following compounds.

-List how many unshared pairs of electrons are in the molecule.

20. HBr

$$\begin{array}{r}
 1+7=8 \\
 -2 \\
 \hline
 6 \\
 -6 \\
 \hline
 0
 \end{array}$$



unshared pairs: 3

21. NI₃

$$\begin{array}{r}
 5+3(7)=26 \\
 -6 \\
 \hline
 20 \\
 -18 \\
 \hline
 2 \\
 -2 \\
 \hline
 0
 \end{array}$$



unshared pairs: 10

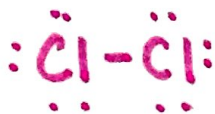
22. N₂



unshared pairs: 2

23. Cl₂

14



unshared pairs: 6

24. SO₂

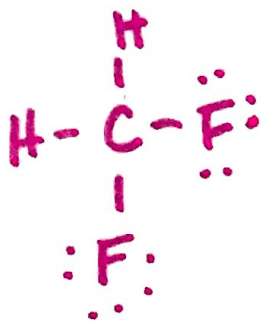
$$\begin{array}{r}
 18 \\
 -4 \\
 \hline
 14
 \end{array}$$



unshared pairs: 6

25. CH₂F₂

$$4+2+14=20$$



unshared pairs: 6