KEY

STEM Chem Semester 2 Unit1 Test Review

ALT 1: I can describe the structure of an atom

AST1.4: I can describe how models of the structure of atoms have changed over time.
AST 1.7: I can describe the electron configurations of a given atom
AST 1.8: I can describe the process by which an atom can emit light.
AST 1.9: I can explain why different atoms have unique emission/absorption spectra.

ALT 2: I can describe the organization of the Periodic Table, find patterns, and predict properties of elements.

AST 2.5: I can identify the groups/regions of the Periodic Table with common characteristics, and can predict properties of elements.

AST 2.6: I can explain the trends in the Periodic Table, including ionization energy and electronegativity.

ALT 5: I can describe how and why atoms form bonds.

AST 5.4: I can compare the characteristics of ionic, covalent, and metallic compounds.
AST 5.5: I can use models to explain the role of valence electrons in bond formation.
AST 5.6: I can describe the relationship between electronegativity and bond type.
AST 5.7: I can use the VESPR model to explain molecular geometry.
AST 5.8: I can compare/contrast the different types of intermolecular forces
AST 5.9: I can use molecular structure to predict intermolecular forces.

Determine the identity of the following elements:

1) I am a transition metal with the smallest atomic radius in my group, I have 6 valence electrons.

~~F~~ Cr

2) I am a non-metal and I belong to the halogen family. I am not the largest or smallest atom in my group, my first ionization energy is greater than that of iodine. I am not a gas at room temperature.

Br

3) I am a main group element with 3 p-orbital electrons in my valence shell. I have a lower first ionization energy than arsenic and I am not the larges atom in my group.

Sb

4) The radius of my most common ion is smaller than my atomic radius, my valence shell contains only one full s-orbital of electrons. I have a lower first ionization energy than calcium and a smaller atomic radius than barium.

Sr

5) The radius of my most common ion is larger than my atomic radius, I have 4 p-orbital electrons in my valence shell. I have a higher first ionization energy than tellurium and I am the smallest atom in my group.

O

Multiple Choice
Identify the choice that best completes the statement or answers the question.

____ 1. Mendeleev noticed that properties of elements usually repeated at regular intervals when the elements were arranged in order of increasing
a. atomic number.
b. density.
c. reactivity.
d. ~~atomic mass.~~

____ 2. Mendeleev predicted that the spaces in his periodic table represented
a. isotopes.
b. radioactive elements.
c. permanent gaps.
d. ~~undiscovered elements.~~

____ 3. Argon, krypton, and xenon are
a. alkaline earth metals.
b. noble gases.
c. actinides.
d. lanthanides.

____ 4. In the modern periodic table, elements are ordered according to
a. decreasing atomic mass.
b. Mendeleev's original design.
c. increasing atomic number.
d. the date of their discovery.

____ 5. The atomic number of lithium, the first element in Group 1, is 3. The atomic number of the second element in this group is
a. 4.
b. 10.
c. 11.
d. 18.

____ 6. How many elements are in a period in which only the *s* and *p* sublevels are filled?
a. 2
b. 8 ← strict
c. 18 ~ sort of
d. 32

____ 7. Period 4 contains 18 elements. How many of these elements have electrons in the *d* sublevel?
a. 8
b. 10
c. 16
d. 18

____ 8. Calcium, atomic number 20, has the electron configuration [Ar] $4s^2$. In what period is calcium?
a. Period 2
b. Period 4
c. Period 8
d. Period 20

____ 9. Elements to the right side of the periodic table (*p*-block elements) have properties most associated with
a. gases.
b. nonmetals.
c. metals.
d. metalloids.

____ 10. Neutral atoms with an s^2p^6 electron configuration in the highest energy level are best classified as
a. metalloids.
b. metals.
c. nonmetals.
d. gases.

____ 11. Which orbitals are characteristic of the lanthanide elements?
a. *d* orbitals
b. *s* orbitals
c. *f* orbitals
d. *p* orbitals

Name: ______________________ **ID: A**

____ 24. A covalent bond results when ____ are shared.
a. ions
b. Lewis structures
c. electrons
d. dipoles

____ 25. The bond type for the C—F bond (electronegativity for C is 2.5; electronegativity for F is 4.0) in CF_4 is
a. ionic
b. non-polar covalent
c. polar covalent
d. hydrogen

____ 26. The elements of the ____ group satisfy the octet rule without forming compounds.
a. main
b. noble gas
c. alkali metal
d. alkaline-earth metal

____ 27. When the octet rule is satisfied, the outermost ____ are filled.
a. *d* and *f* orbitals
b. *s* and *p* orbitals
c. *s* and *d* orbitals
d. *d* and *p* orbitals

____ 28. In drawing a Lewis structure, the central atom is the
a. atom with the greatest mass.
b. atom with the highest atomic number.
c. atom with the fewest electrons.
d. least electronegative atom.

OK if

____ 29. To draw a Lewis structure, one must know the
a. number of valence electrons in each atom.
b. atomic mass of each atom.
c. bond length of each atom.
d. ionization energy of each atom.

____ 30. Multiple covalent bonds may occur in atoms that contain carbon, nitrogen, or
a. chlorine.
b. hydrogen.
c. oxygen.
d. helium.

____ 31. A formula that shows the types and numbers of atoms combined in a single molecule is called a(n)
a. molecular formula.
b. ionic formula.
c. Lewis structure.
d. covalent formula.

32. Describe the process by which an atom can emit light

33. Why do different elements have unique emission spectra?

Electron Configuration Problem Set

Element	# of e^-	Electron Configuration	Previous Nobel Gas	Shorthand Notation	Orbital Diagram
	8		He	[He] $2s^2 2p^4$	[He] $\underset{2s}{\underline{\uparrow\downarrow}}$ $\underset{2p}{\underline{\uparrow\downarrow}\,\underline{\uparrow}\,\underline{\uparrow}}$
		$1s^2 2s^2 2p^6 3s^1$			
Magnesium			Ne		
				[Ar] $4s^2 3d^8$	
	47				
Tin					
		$1s^2 2s^2 2p^3$			
Bromine			Ar		
					[Ne] $\underset{3s}{\underline{\uparrow\downarrow}}$ $\underset{3p}{\underline{\uparrow}\,\underline{\uparrow}\,\underline{}}$
	32		Ar		

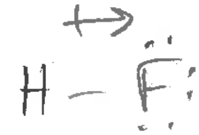
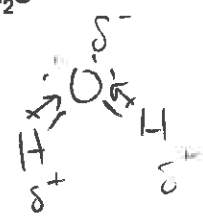
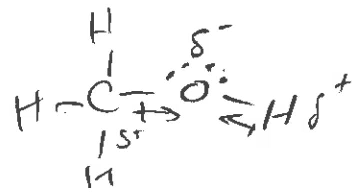
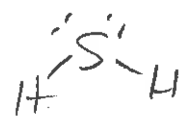
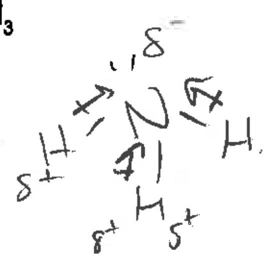
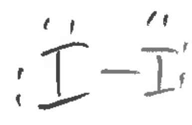
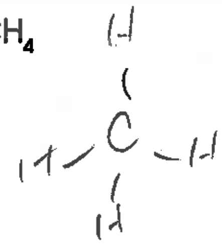
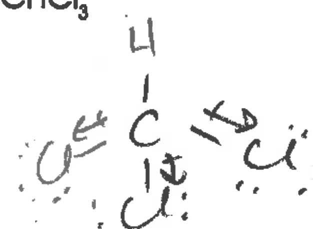
Questions:

1) How many electrons can fit in the 3d sublevel?
2) How many electrons can fit in a 2p orbital?
3) What is wrong with the following orbital diagram? [He] $\underset{2s}{\underline{\uparrow\downarrow}}$ $\underset{2p}{\underline{\uparrow\downarrow}\,\underline{\uparrow\downarrow}\,\underline{}}$
4) Following the pattern for the quantum model, how many sublevels would you expect in the 6^{th} principal energy level?
5) What is wrong with the following electron configuration? $1s^2 2s^1 2p^6 3s^1$

POLARITY OF MOLECULES

Name ______________

Determine whether the following molecules are polar or nonpolar. (Show bond dipoles & partial charges)

1. N_2  non polar	7. HF  polar
2. H_2O  polar	8. CH_3OH  polar
3. CO_2  non-polar	9. H_2S  non polar
4. NH_3  polar	10. I_2  non polar
5. CH_4  non-polar	11. $CHCl_3$  polar
6. SO_3 non polar	12. O_2  non polar