PROBLEM SESSION KEY FOR EXAM 3

1. Breathalyzers estimate the amount of alcohol in the blood by measuring the alcohol in the breath. The breathalyzer uses the redox reaction below to determine the amount of alcohol in the blood. Answer the following questions using this balanced chemical equation.

3 C2H5OH + 10 H2CrO4 🡪 3 CH3CO2H + 8 Cr2(CrO4)3 + 13 H2O + 2674 kJ

ethanol chromic acid acetic acid chromium(III) chromate water

* 1. How many moles of chromic acid can react with 5.92 moles of ethanol?
  2. How many mg of ethanol are in a sample of a driver’s breath that produces 21.8 mg of acetic acid?
  3. How much energy will be produced if 38.1 grams of chromic acid react with excess ethanol?
  4. If 62.8 grams of ethanol react with excess chromic acid to produce 89.4 grams of water, what is the percent yield of the reaction?
  5. If 9.25 g of ethanol are allowed to react with 62.8 g or chromic acid, how many g of acetic acid should be produced?

Chromic acid is limiting and only 9.59 g of acetic acid will be produced.

1. Write a balanced equation for the reaction, if any that occurs in each of the following cases. Assume that all soluble reactants are added in the form of aqueous solutions. Indicate gases and precipitates that are formed, as well as insoluble solid reactants. If no reaction occurs, then write **NO RXN**, and do not write a balanced equation. Be sure to **balance** your equations and include your **phase labels**.
2. potassium hydroxide + sodium hydroxide 🡪

KOH (aq) + NaOH (aq) 🡪 KOH (aq) + NaOH (aq) 🡪 no reaction

1. sodium acetate + hydrochloric acid 🡪

NaC2H3O2 (aq) + HCl (aq) 🡪 NaCl (aq) + HC2H­3O2 (aq)

1. zinc bromide + potassium phosphate 🡪

3 ZnBr2 (aq) + 2 K3PO4 (aq) 🡪 6 KBr (aq) + Zn3(PO4)2 (s)

1. hydrochloric acid + calcium 🡪

2 HCl (aq) + Ca (s) 🡪 CaCl2 (aq) + H2 (g)

1. nitric acid + barium hydroxide 🡪

2 HNO3 (aq) + Ba(OH)2 (aq) 🡪 Ba(NO3)2 (aq) + 2 H2O (l)

1. ammonium nitrate + sodium hydroxide 🡪

NH4NO3 (aq) + NaOH (aq) 🡪 NaNO3 (aq) + NH3 (g) + H2O (l)

1. How many electrons are there in an orbital? \_\_\_2 e-\_\_\_\_\_\_

In an s sublevel? \_\_\_2 e-\_\_\_\_\_

A p sublevel?\_\_\_\_ 6 e-\_\_\_\_\_

A d sublevel?\_\_\_\_ 10 e-\_\_\_\_\_

1. How is the line spectra of an atom produced? (Answer this on an atomic level)

Electrons bound to atoms are in quantized energy levels (orbits have specific energy with most energies NOT allowed). The energy levels are discrete (quantized - as in quantum mechanics). These electrons only have certain choices for energy levels (orbits). This means that when an excited electron gives off energy it gives off exactly the amount which is the difference between the orbits it is on the orbit it is allowed to "drop" into. Since there are rules which make some destination orbits more likely than others, each atom has a characteristic emission spectra, certain light frequencies are emitted more than others (so are brighter) giving each element a unique spectra.

1. How do atomic orbitals fill? If a p sublevel has 4 electrons, which orbitals will they occupy? Draw the sublevel using arrows to represent electrons and show spin based on the direction of the arrow.

How do atomic orbitals fill? Electrons fill energy levels and orbitals from the lowest energy to the highest energy.

1. the number of electrons in an atom is equal to the atomic number;
2. each added electron will enter the orbitals in the order of increasing energy;
3. an orbital cannot take more than 2 electrons.

**Pauli's Exclusion Principle:** Each orbital can hold two electrons with opposite spins

**Hund's Rule** Electrons enter orbitals of equal energy one at a time with parallel spin to minimize repulsion only when each equal energy orbital contains in electron do you begin to pair the electrons gives those opposite spin.

(From question) If a p sublevel has 4 electrons, which orbitals will they occupy? Draw the sublevel using arrows to represent electrons and show spin based on the direction of the arrow.



1. Explain why each of the following electron configurations for a p sublevel is disallowed.
   1. violates Pauli's Exclusion Principle, 1st box the electrons need to have opposite spins
   2. violatesHund's Rule, an electron needs to be placed in each equal energy orbital with parallel spin before pairing the electrons
   3. An orbital can only hold 2 electrons, 1st box contains 3 electrons
   4. violatesHund's Rule, an electron needs to be placed in each equal energy orbital with parallel spin. The second box the electron does not have parallel spin to the other two electrons.
2. Write the complete and shorthand electronic configuration for the following atoms and ions.

|  |  |
| --- | --- |
| shorthand | complete |
| Si: **[Ne] 3s2 3p2** | Si: 1s22s22p63s23p2 |
| Mg+2 : [He] 2s22p6 | Mg+2 : 1s22s22p6 |
| S‑2: [Ne] 3s23p6 | S‑2: 1s22s22p63s23p6 |
| V: **[Ar] 3d3 4s2** | V: 1s22s22p63s23p64s23d3 |
| Mn: **[Ar] 3d5 4s2** | Mn: 1s22s22p63s23p64s23d5 |
| Ni +2: **[Ar] 3d8** | Ni +2: 1s22s22p63s23p63d8 |

1. Explain why an atom might have an anomalous configuration. Predict which elements might have anomalous configurations.

Reason for this anomalous electron configuration - half-filled and completely filled sets of orbitals provide extra stability and lower energy to the atom.

can reduce electron-electron repulsions,

the orbital energies of elements lower in the periodic table start becoming very similar, so electrons are freer to move around between different orbitals than in the lighter elements.

In chromium, lower energy is achieved by half-filling the 4s orbital and all five 3d orbitals. There is greater electron-electron repulsion by putting two electrons in the 4s orbital when there is room available in an unoccupied 3d orbital, so the half-filled subshell arrangement is preferred.   
  
Chromium and copper are the most important exceptions in the first row of transition metals. As stated, chromium is 4s1, 3d5. Copper has the configuration 4s1, 3d10, because again, the full set of d orbitals provides extra stability.   
  
You will find anomalous electron configurations as you go farther down into the transition metals. The elements below chromium and copper follow similar patterns with half-filled and completely filled d orbitals. However, there are also other exceptions scattered throughout the heavier elements as d orbital and f orbital energies become increasingly similar

1. How many valence electrons in an atom of phosphorous? Of barium? Of krypton?

P: 5 valence electrons Ba: 2 valence electrons Kr : 8 valence electrons

1. Explain how an ionic bond differs from a covalent bond.

In ionic bonding, atoms transfer electrons to each other. Ionic bonds require at least one electron donor and one electron acceptor (usually metal + nonmetal). In contrast, atoms that have the same electronegativity share electrons in covalent bonds since donating or receiving electrons is unfavorable (usually nonmetal + nonmetal)..

1. Explain how a pure covalent bond differs from a polar covalent bond.

Pure Covalent Bond

* A bond between 2 nonmetal atoms that have the same electronegativity and therefore have equal sharing of the bonding electron pair

Polar Covalent Bond

* A bond between 2 nonmetal atoms that have different electronegativity’s and therefore have unequal sharing of the bonding electron pair

1. What is the lewis electron dot structure of carbon? Of arsenic? Of potassium?

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1. Rank the following atoms in order of increasing radius:

2 N 3 Sb 4 Rb 5 Cs 1 F

1. Name the element that corresponds to each of the following:
2. alkali metal with the smallest atomic radius Li
3. Group 5A element with the highest ionization energy N
4. [Kr] 5s24d10 Cd

1. Give three anions that are isoelectronic with neon?

F-1, O-2, N-3

1. List the following in increasing (L to R) ionization energy.

N, Ca, Cl, Fr, Rb

Fr < Rb < Ca < Cl < N

Increasing Ionization Energy

1. Arrange the following elements in order of increasing atomic size: Ca, B, S, Si, Ge, F

F< B<S<Si<Ge< Ca

1. Arrange the following elements in order of increasing metallic character: Fr. Sb. In, S, Ba, Se

S<Se<Sb<In<Ba<Fr

1. Draw Lewis Electron Dot Structures for the following molecules and tell the orbital and molecular geometries for each

|  |  |
| --- | --- |
| * 1. PH3 Val e- : 8     Orbital geometry: tetrahedral  Molecular geometry: trigonal pyramidal | * 1. C2H4 Val e- : 12     Orbital geometry: Trigonal planar  Molecular geometry: Trigonal planar |
| * 1. CBr4 Val e- : 32     Orbital geometry: tetrahedral  Molecular geometry: tetrahedral | * 1. C2H6 Val e- : 14     Orbital geometry: tetrahedral  Molecular geometry: tetrahedral |
| * 1. OF2 Val e- : 20     Orbital geometry: tetrahedral  Molecular geometry: bent | * 1. N2H2 Val e- : 12     Orbital geometry: Trigonal planar  Molecular geometry: bent |
| * 1. SCl2 Val e- : 20     Orbital geometry: tetrahedral  Molecular geometry: bent | * 1. C2H2 Val e- : 10     Orbital geometry: linear  Molecular geometry: linear |
| * 1. CO2 (C is central) Val e- : 16     Orbital geometry: linear  Molecular geometry: linear | * 1. Cl2CO (C is central) Val e- : 24     Orbital geometry: Trigonal planar  Molecular geometry: Trigonal planar |

1. Draw Lewis Electron Dot Structures for the following ions. Include resonance structures as necessary.

|  |  |
| --- | --- |
| * 1. CN-1 Val e- : 10 | * 1. NO2-1 Val e- : 18 |
| * 1. SO3-2 Val e- : 26 | * 1. CO3-2 Val e- : 24 |
| * 1. NH4+1 Val e- : 8 | * 1. ClO3- Val e- : 26 |

1. Predict the orbital and molecular geometry around each of the starred atoms in the drawing below:



|  |  |  |
| --- | --- | --- |
| Atom | Orbital geometry | Molecular geometry |
| 1 | Trigonal planar | Trigonal planar |
| 2 | tetrahedral | tetrahedral |
| 3 | Trigonal planar | bent |
| 4 | tetrahedral | Trigonal pyramidal |
| 5 | tetrahedral | tetrahedral |
| 6 | linear | linear |
| 7 | Trigonal planar | bent |
| 8 | tetrahedral | bent |