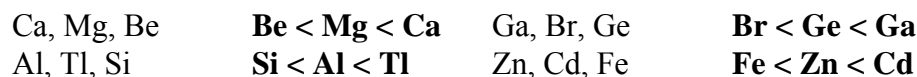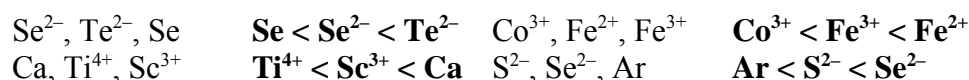**Chem 115 POGIL Worksheet - Week 10 - Solutions**
More on Configurations and Periodic Trends

Key Questions

1. Using only a periodic table, predict the order of increasing atomic radius for each of the following sets of elements:

Ca, Mg, Be	**Be < Mg < Ca**	Ga, Br, Ge	**Br < Ge < Ga**
Al, Tl, Si	**Si < Al < Tl**	Zn, Cd, Fe	**Fe < Zn < Cd**

2. For each of the following sets of atoms and ions, arrange the members in order of increasing size.

Se^{2-}, Te^{2-}, Se	**Se < Se^{2-} < Te^{2-}**	Co^{3+}, Fe^{2+}, Fe^{3+}	**Co^{3+} < Fe^{3+} < Fe^{2+}**
Ca, Ti^{4+}, Sc^{3+}	**Ti^{4+} < Sc^{3+} < Ca**	S^{2-}, Se^{2-}, Ar	**Ar < S^{2-} < Se^{2-}**

3. Based on position in the periodic table, which element of the following pairs has the higher first ionization energy?

    O, **Ne**   **Ca**, Sr   K, **Cr**   **Br**, Sb   In, **Sn**

4. First ionization energies tend to increase across a period. But in period 2, Be has a higher first ionization energy than B, and N has a higher first ionization energy than O. Explain, [Hint: Look at the valence configurations, and recall that half-filled and fully filled subshells have extra stability.]

    Compare the valence configurations of the atoms and the ions formed by removing one electron:

Element	X	X^+
Be	$2s^2$	$2s^1$
B	$2s^22p^1$	$2s^2$
N	$2s^22p^3$	$2s^22p^2$
O	$2s^22p^4$	$2s^22p^3$

    Be has a stable closed subshell configuration, which is disrupted by losing an electron. By contrast, B can achieve a fully-filled subshell configuration by losing the lone electron in the $2p$ subshell. This makes ionization of Be a little higher than usual, and ionization of B a little lower than usual. This causes the reversal of trend (the "jog" in the plot) across these two elements. A similar thing occurs between N and O. N has a stable half-filled $2p$ subshell, and O can achieve a stable half-filled subshell by losing an electron. Thus, the ionization energy of N is higher than O, resulting in a "jog" in the plot of ionization energies across these two elements.

5. On the basis of electronic configurations or any other appropriate considerations, explain the differences in the electron affinities of the following pairs of species.

$Na, A_1 < 0; Mg, A_1 > 0$ $N, A_1 > 0; O, A_1 < 0$
$Br, A_1 \ll 0; Kr, A_1 > 0$ $O, A_1 < 0; O^-, A_2 > 0$

$Na, A_1 < 0; Mg, A_1 > 0$ Adding an electron to Na results in a stable closed-subshell $3s^2$ configuration, but with Mg it goes beyond this to $3p^1$, which has no special stability.

$N, A_1 > 0; O, A_1 < 0$ The additional electron disrupts the stable $2p^3$ half-fill subshell configuration of N. The $2p^4$ configuration of O has no special stability, but the higher nuclear charge of O makes acquiring the extra electron favorable.

$Br, A_1 \ll 0; Kr, A_1 > 0$ When Br acquires an electron it completes its $4p$ subshell, resulting in a stable $4s^2 4p^6$ closed-subshell configuration. Kr already has that configuration, and an extra electron must be added to the next shell, resulting in the configuration $5s^1$.

$O, A_1 < 0; O^-, A_2 > 0$ The first electron affinity of oxygen is negative because of the small size and relatively high nuclear charge. Once O^- is formed, the next electron must overcome charge repulsions to be added, which makes the second electron affinity unfavorable and positive.
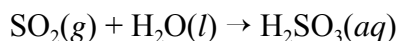
6. Calcium is generally less reactive than potassium but more reactive than magnesium. Explain.

Potassium has a lower ionization energy, and only a single electron loss is needed to make its stable ion, K^+. Calcium must lose two electron to form its stable ion, Ca^{2+}, and the energy cost of the two ionizations is vastly greater than the single ionization of potassium. The two ionizations of magnesium to form Mg^{2+} are even higher, because magnesium is smaller than calcium. Therefore, magnesium is the least reactive of the three.
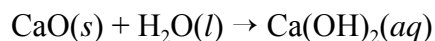
7. For each of the following oxides, indicate whether it is ionic or molecular and whether it is acidic or basic. Then, write a balanced equation for the reaction expected between each oxide and water.
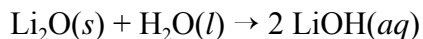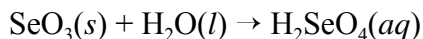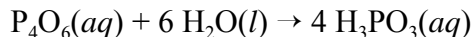
$$SO_2(g),\ CaO(s),\ Li_2O(s),\ SeO_3(s),\ P_4O_6(s)$$

$SO_2(g)$ – molecular, acidic
$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$
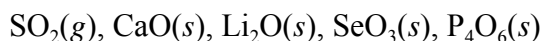
$CaO(s)$ – ionic, basic
$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$$

$Li_2O(s)$ – ionic, basic

$$Li_2O(s) + H_2O(l) \rightarrow 2\ LiOH(aq)$$

$SeO_3(s)$ – molecular, acidic

$$SeO_3(s) + H_2O(l) \rightarrow H_2SeO_4(aq)$$

$P_4O_6(s)$ – molecular, acidic

$$P_4O_6(aq) + 6\ H_2O(l) \rightarrow 4\ H_3PO_3(aq)$$

8. Using the electronegativity values in the table above, calculate $\delta\chi$ for the pairs of elements forming bonds in each of the following species, and describe the type of bonding as either pure covalent, polar covalent, or ionic: H_2, HCl, $PbBr_2$, NaCl, NCl_3.

H_2	$\delta\chi = \|2.1 - 2.1\| = 0$	pure covalent
HCl	$\delta\chi = \|2.1 - 3.0\| = 0.9$	polar covalent
$PbBr_2$	$\delta\chi = \|1.8 - 2.8\| = 1.0$	polar covalent
NBr_3	$\delta\chi = \|3.0 - 2.8\| = 0.2$	polar covalent (slightly)
NaCl	$\delta\chi = \|0.9 - 3.0\| = 2.1$	ionic

9. Calculate $\delta\chi$ for the bonds in HF and NCl_3 and describe them.

HF	$\delta\chi = \|2.1 - 4.0\| = 1.9$	very polar covalent

Although $\delta\chi$ suggests an ionic bond, HF is a gas and exists as a molecular species. Its bonds are very polar covalent, but HF does not contain discrete H^+ and F^- ions like a typical ionic compound.

NCl_3	$\delta\chi = \|3.0 - 3.0\| = 0$	virtually pure covalent

The zero value of difference in electronegativity suggests that the polarity is very low in the N–Cl bonds in NCl_3. But because N and Cl are two different elements, there is actually some difference in their abilities to attract electrons in the bond. Electronegativity is simply too imprecise to indicate the polarity.

10. Explain the following trends in lattice energy.
   a. KF < CaO < ScN
      The ion charges are +1/–1 for KCl, +2/–2 for CaO, and +3/–3 for ScN. Lattice energy rises sharply with increasing ion charge.
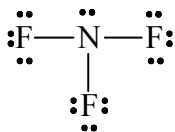
   b. NaCl > RbBr > CsBr
      All of these are +1/–1 ion pairings, so size is the dominating factor. As ion size increases, the separation between oppositely charged ions increases, causing the forces of attraction to diminish.
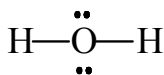
c. $MgO > MgCl_2$
The Cl^- ion is larger and has a lesser charge than the O^{2-} ion. Both factors result in $MgCl_2$ having a lower lattice energy than MgO.

11. Following the steps given above, draw Lewis dot structures for the following molecular species: NF_3, H_2O, NH_4^+, SO_4^{2-}, H_2CO, C_2H_2.
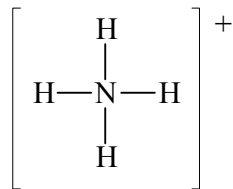
NF_3 $N + 3F = 5 + (3)(7) = 26$

$$:\ddot{F}\!-\!\ddot{N}\!-\!\ddot{F}:$$
$$:\ddot{F}:$$

H_2O $O + 2H = 6 + (2)(1) = 8$

$$H\!-\!\ddot{O}\!-\!H$$

NH_4^+ $N + 4H + (+) = 5 + (4)(1) - 1$

$$\left[H\!-\!\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{N}}\!-\!H \right]^+$$

SO_4^{2-} $S + 4O + (2-) = 6 + 4(6) + 2 = 32$

$$\left[\ddot{:O}: \\ :\ddot{O}\!-\!\underset{\underset{:\ddot{O}:}{|}}{\overset{\overset{|}{}}{S}}\!-\!\ddot{O}: \right]^{2-}$$

H_2CO $2H + C + O = (2)(1) + 4 + 6 = 12$

$$:O: \\ \| \\ H\!-\!C\!-\!H$$

C_2H_2 $2C + 2H = (2)(4) + (2)(1) = 10$

$$H\!-\!C\!\equiv\!C\!-\!H$$
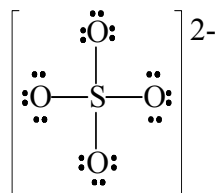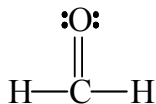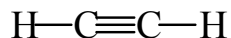
12. Following the steps above, draw Lewis dot structures for the following molecular species, which may be electron deficient or hypervalent: $AlCl_3$, XeF_4, XeF_2, $BeCl_2$

$AlCl_3$ $Al + 3Cl = 3 + (3)(7) = 24$

:Cl—Al—Cl: Al is electron deficient (3 pairs).
      |
    :Cl:

XeF_4 $Xe + 4F = 8 + (4)(7) = 36$

Xe is hypervalent (6 pairs).

:F:
 |
:F—Xe—F:
 |
:F:

XeF_2 $Xe + 2F = 8 + (2)(7) = 22$

:F—Xe—F: Xe is hypervalent (5 pairs).

$BeCl_2$ $Be + 2Cl = 2 + (2)(7) = 16$

:Cl—Be—Cl: Be is electron deficient (2 pairs)
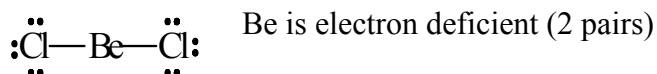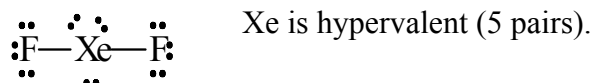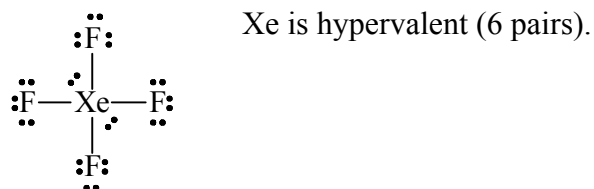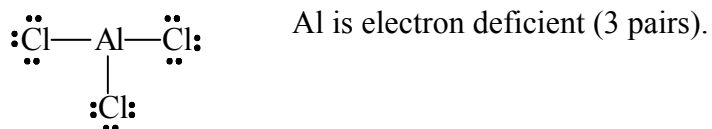
13. Show the formal charges on all atoms in the sulfate ion, SO_4^{2-}, whose Lewis dot structure you developed in Key Question 4.
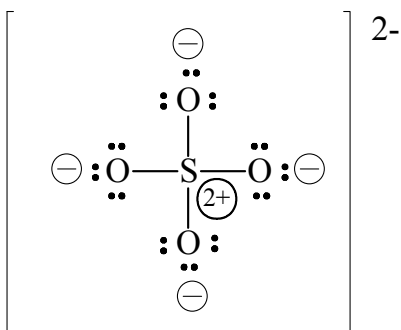
14. The nitrate ion, NO_3^-, can be represented by a series of three equivalent resonance forms. Show the resonance forms of NO_3^-, including formal charges, and estimate the bond order of the N–O bonds in this ion.

NO_3^- $N + 3O + (-) = 5 + (3)(6) + 1 = 24$
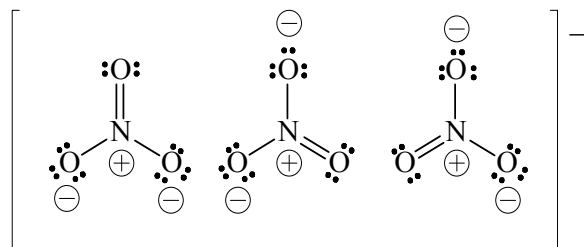
As with ozone, the single and double bonds in these structures do not really exist in the nitrate ion. All N-O bonds are equivalent. The three N-O bonds in any one of the three equivalent resonance forms involve four pairs of electrons. Therefore the approximate bond order for a N-O bond in the nitrate ion is $4/3 = 1⅓$.

15. FNO_2 can be represented by three non-equivalent resonance forms. Draw the three canonical forms of FNO_2 and assess their relative contributions to the resonance hybrid description of the molecule. What is the approximate N–O bond order?

FNO_2 $F + N + 2O = 7 + 5 + (2)(6) = 24$
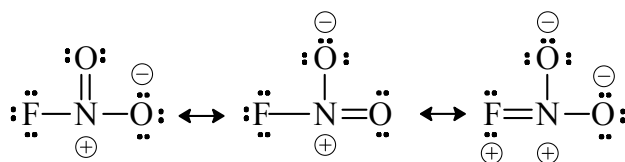
The third form is insignificant (adjacent like formal charges). In the remaining forms, the two N–O bonds are approximately equivalent. The two N–O bonds use 3 pairs, so the bond order of each is about $3/2 = 1½$.