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Chapter 6 Practice Worksheet:
Ionic Bonds and Some Main-Group Chemistry

1) Write electron configurations of the following ions:

a) K^+ (long hand notation): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

b) F^- (long hand): $1s^2\ 2s^2\ 2p^6$

c) Ni^{2+} (short hand): $[Ar]\ 4s^0\ 3d^8$

d) Zn^{2+} (short hand): $[Ar]\ 4s^0\ 3d^{10}$

e) Br^- (short hand): $[Ar]\ 4s^2\ 3d^{10}\ 4p^6$

f) Cu^{2+} (short hand): $[Ar]\ 4s^0\ 3d^9$

g) Ag^+ (short hand): $[Kr]\ 5s^0\ 4d^{10}$

2) Draw orbital diagrams (for valence electrons only) for the following ions:

a) K^+ $\underline{\uparrow\downarrow}$ (3s) $\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}$ (3p)

b) Br^- $\underline{\uparrow\downarrow}$ (4s) $\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}$ (3d) $\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}$ (4p)

c) Ni^{2+} ___ (4s) $\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\ }\ \underline{\uparrow\ }$ (3d)

d) Zn^{2+} ___ (4s) $\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}$ (3d)

3) Define the following terms: **atomic radius** and **ionic radius**. Describe the periodic trend for each one.

Atomic radius measures half the distance between the nuclei of two adjacent atoms. (Since the orbital boundaries of atoms are not absolutely defined, atomic radius cannot be measured for a single atom.) Ionic radius is defined similarly but for ions in an ionic compound. Both periodic trends increase going down a group. Atomic radius generally decreases going from left to right across the periodic table. Ionic radius also decreases left to right but jumps in size at the transition between cations and anions.

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4) Define Z_{eff} (effective nuclear charge). How does this property affect radii of atoms as you move from left to right in a period? Does it affect atomic radius as you move down a group?

Effective nuclear charge defines strength of the attraction between protons and electrons. As you move from left to right in across the periodic table, the number of protons increases, but the valence electrons are in the same shell.

5) Arrange the following atoms in order of **increasing** radius: N, Sb, P, Bi, As

N < P < As < Sb < Bi

6) Define the phrase **isoelectronic series**. Describe how sizes of ions change as you move from the most positively charged ion to the most negatively charged ion in an isoelectronic series.

Isoelectronic means ions that have the same electron configuration. Sizes of ions increase as you move from the most positively charged ion to the most negatively charged ion in a series.
Example: S^{2-} has 18 e^- and 16 p^+; Ca^{2+} has 18 e^- and 20 p^+. Ca has more protons to pull in its electrons and results in a smaller ion.

7) Arrange the following ions in order of **increasing** radius: F^-, Na^+, O^{2-}, Mg^{2+}, N^{3-}

$Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$

8) Define: **ionization energy** (write a chemical equation to demonstrate this) and **electron affinity** (write a chemical equation to demonstrate this).

Ionization energy is the amount of energy required to remove an electron from an element: $El \rightarrow El^+ + e^-$
Electron affinity is the amount of energy given off when an element gains an electron: $El + e^- \rightarrow El^-$

9) Show the direction in which each trend from #8 increases. Use the periodic table below to draw arrows of increasing value for each property.
Both trends increase going up a group and from left to right across the periodic table.

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Ha	106 Sg	107 Ns	108 Hs	109 Mt	110	111	112	(113)	(114)	(115)	(116)	(117)	(118)

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10) Arrange the following atoms in order of **increasing** first ionization energy: Ba, Ca, Be, Sr, Mg

Ba < Sr < Ca < Mg < Be

11) Arrange the following atoms in order of **increasing** electron affinity: Br, Sb, I, Te, Cl

Sb < Te < I < Br < Cl

12) What two properties affect the lattice energy of an ionic compound? What is the relationship between these properties and lattice energy (i.e., are they inversely or directly related to lattice energy?)?

Charge of ion (directly related to lattice energy); Radius (inversely related to lattice energy)

13) What physical properties does lattice energy affect?
Melting point, boiling point, heats of fusion and vaporization

14) Does a stronger or weaker lattice energy result in a stronger ionic bond? How will this affect melting and boiling points of a crystal lattice?

Stronger lattice energy results in a stronger bond. The stronger the bond, the more energy (as heat) required to separate ions.

15) Determine which compound in each pair below will have a higher lattice energy. Defend your answer on the line to the right.

a. $NaCl$ $NaBr$ __Cl is smaller than Br (all ions have a charge of +/-1__________

b. KF CaF_2 __Ca is smaller and has a larger charge than K_________________

c. MgO Na_2O __Mg is smaller and has a larger charge than Na________________

d. KF $CsCl$ __K and F are smaller than both Cs and Cl (same charges on all)__

e. $RbBr$ $CaCl_2$ Ca is smaller and has a larger charge than Rb; Cl is smaller than Br