WS 1: Ionic Bonds Answer Key[1]

Part I: The Ionic Bonding Model, Electrostatic attractions:[i]

Questions to Help You Think About the Concept

1 If the charges (Q_1 & Q_2) remain constant, what happens to the magnitude of the electrostatic energy if the particles are moved farther apart?

The magnitude of the energy decreases because energy is inversely proportional to d, the distance between the charges.

2 If the particles are moved so far apart that they no longer feel an attractive force, what happens to the magnitude or value of the ΔE_{el}?

The value of the energy is zero because the attractive "force" has been reduced to zero.

3 If d is a small, finite value, and $Q_1 = Q_2$, is the $|\Delta E_{el}|$ a large value or a small value?

The absolute value of the energy would be a larger number than if the distance was large. Since the charges are the same, the main effect on the energy would be the distance between the charges or nucli.

4 According to the electrostatic equation: if a particle with a charge of –1 is really, really close to a particle with a charge of 2+, the potential energy to separate the ions is: a) large and positive, b)large and negative, c)small and negative, or d) small and positive.

If the ions are close to each other and the charge is increasing (from maybe +1 to +2), the attraction between the ions is increasing. It would take lots of energy to separate these ions. The energy would be large, and it would be positive.

Part II: Ionic Bonds and Coulomb' Law

Questions To Start You Thinking:

1. In LiCl, what are the charges on the ions?
   The charge of the lithium ion is 1+, and the chloride ion is 1–

2. Looking at the picture of sodium chloride in your textbook or on the internet; the two different types of ions are represented by spheres of different sizes.
   a) Which spheres represent the sodium ions and which spheres represent the chloride ions? **The larger spheres represent the chloride ion, while the smaller sphere represents the sodium ion. This makes sense to me because when a cation is created, an electron is removed, reducing the shielding. The radius decreases. when an electron is added, the shielding increases, and the radius**

[1] This information is covered in Brown and LeMay sections 8.2-8.3.

of the ion increases because attraction to the outermost electron has decreased.

3. Now imagine the compound magnesium oxide.

   a) What is the charge on the cation formed from Mg?
   b) What is the charge on the anion formed from O?
   c) What is the formula for the ionic compound containing magnesium ions and oxide ions?
   d) Which compound should have the largest lattice energy, magnesium oxide or sodium chloride?

The charge on the magnesium is 2+; the charge on the oxide ion is 2–. The formula is MgO. Magnesium oxide will have the largest lattice energy because it has the largest attraction between the two ions. The size of the magnesium ion is smaller than the sodium ion; the size of the oxide ion (though having a larger charge) is smaller than chloride because they are in different '*n*' shells. A smaller distance coupled with a larger charge equates to a larger lattice energy.

4. Which is larger a fluoride ion or a chloride ion? Explain your answer.

The chloride ion is in a larger '*n*' shell. The electron is further from the nucleus, the electrostatic attraction should be lower, and the radius should be larger.

5. Account for the following variation in lattice energies (in kJ/mol) observed for the potassium halides:

KF [817], KCl [718], KBr [688], KI [636].

The anions in this series are in the same family. size increases down a family. as the ions are further apart, the attraction between the ions decreases while the charge stays constant. The melting point, which is a measure of the energy needed to break the attraction between the ions, is decreases because less energy is needed to affect this change.

Problems and Exploration:

1. When Na(s) and Cl_2(g) react, the ionic compound NaCl(s) forms. Predict the formula of ionic compounds formed from the following combinations of elements:

   a) Sodium metal and liquid bromine **NaBr**
   b) Lithium metal and oxygen gas **Li_2O**
   c) Aluminum metal and nitrogen gas **Al_2N_3**
   d) Magnesium metal and liquid bromine **$MgBr_2$**
   e) Calcium metal and oxygen gas **CaO**

2. A representative group metal reacts with chlorine and with oxygen to form ionic compounds with the formulas MCl_4 and MO_2. Propose a possible identity for the metal M. Explain your reasoning.

The metal is likely to be either Ge, Sn, or Pb. These are elements in 4A, which can form 4 bonds.

3. One measure of the strength of the bonds holding the ions together is the melting point: the more strongly the ions are held together, the higher the melting point.

The units for ΔE_{el} is kJ/mol, 1e = $1.602x10^{-19}$C, d is the meter, and k = $8.99X10^{9}$ J/m•C^{2-}.
WATCH YOUR UNITS!

Example: Rb_2O we consider the charge between the Rb^+ and the O^{2-}

$$\frac{8.99x10^9 J}{mx\ C^2} x \frac{(1+)(2-)x(1.602\ x10^{-19}C)^2}{1.66\ x10^{-10}m + 1.26\ x10^{-10}m} x \frac{(6.022x10^{23}\ ion\ pair)}{mole} \frac{kJ}{1000J}$$
$$= 476\ kJ/mol$$

Now you try!

Ionic compound	Radius (in pm)		D= $r_{cation} + r_{anion}$ in pm	Charge		$E_{El} = \frac{kQ_1Q_2}{d}$ (kJ/mol)	Melting point
	Cation	anion		Cation	Anion		See
NaF	102	133	235	1+	1–	591	**993°C**
NaCl	102	181	283	1+	1–	**491**	**801°C**
MgO	72	140	212	2+	2–	**$2.62x10^3$**	**2852°C**
MgS	72	184	256	2+	2–	**$2.17x10^3$**	**2000°C**

Fill in the chart above before answering the questions below. Melting points are presented in problems below

Questions To Start You Thinking:

1. Consider the ionic compounds NaF and NaCl. In which compound is the columbic force of attraction greater? NaCl has a melting point of 801°C. Which of these would you predict is the melting point of NaF: 609°C, 800°C, 993°C? Explain your reasoning.

Since the charge is the same for both of these ion pairs, the melting point of NaF must be larger than that of NaCl-993°C.

2. Consider the ionic compounds MgO and MgS: In which compound is the columbic force of attraction greater? MgO has a melting point of 2852°C. Which of these would you predict is the melting point of MgS: ≈2000°C, ≈2850°C, ≈ 4000°C? Explain your reasoning.
3. **Since the charge is the same for both of these ion pairs, the melting point of NaF must be larger than that of NaCl-993°C.**

4. Based on your choices for the melting points, which factor, the size of the ions or the charge of the ions has the larger effect on the melting point?

The larger the charge and the smaller the radii, the larger the melting point. However, the charge has a larger effect than the size of the radii. If I compare NaF (235 pm) with MgS (256pm) the charge of the ions for the MgS has a larger effect on the melting point than that of the NaF because of the charges on the magnesium and the sulfide.

Problems and Exploration:

1. Predict which ionic compound has the higher melting point in each of the following pairs

   a) NaCl and NaBr

   b) NaCl and KCl

   c) MgO and CaO

   d) KCl and CaO

   e) NaCl and MgS

   f) NaCl and NaF

2. Rank the following compounds in order of increasing melting point and explain your reasoning:
   KCl K_2S CaS CaO

3. For each of the following, which compound is expected to have the highest melting point? Explain your choice in one or two sentences.

   a) LiF; LiCl; NaF; NaCl; KI: smallest radii for same charges

   b) NaF; NaCl; CaS; CaO: largest charge for the ion pair, radius of O is smaller than the radius of S (n=2 vs n = 3)

   c) Na_2SO_4; K_2SO_4; $CaSO_4$; $BaSO_4$ largest charge for the ion pair, radii of anion same, Ca smaller than Ba (n= 4 vs n = 6)

   d) LiF; CaO; BaO; Al_2O_3 largest charge for the ion pair, Ba ion too big, lower charge than Al, Ca smaller radius than Ba, but still smaller charge than Al,

   e) H_2O; NH_3; N_2; $CaSO_4$; O_2: nitrogen is a gas as is ammonia and oxygen, so there melting points are very low. The ionic compound would have the highest melting point since it is a solid at room temperature and the others are not (melting belowroom temperature)

Part III: Ionic bonds

Questions To Start You Thinking:

1. Compounds that contain ions formed from group 1A atoms have a 1+ charge, rather than a 2+ charge. Explain this based on Z-effective, electron affinity, ionization energy, and/or electron configuration.

When a group 1A element loses an electron, the shielding goes down. This is the basis of ionization energy, the energy needed to remove an electron in the gas phase. The higher the Z_{eff}, the higher the ionization energy. Removing an electrons raises the Z_{eff}. The

electron configuration created is iso-electronic to a noble gas, but with a higher Z_{eff}. Removing another electron would be difficult because the (new) outer most electron has a stronger attraction to the nucleus. It would be energetically unfavorable to remove the second electron to create the 2+ charge.

2. Compounds that contain ions formed from group 7A atoms are —1. Why would it not be 2–? Explain this based on Z-effective, electron affinity, ionization energy, and/or electron configuration.

When a group 7A element gains an electron, the shielding goes up. Adding an electrons lowers the Z_{eff}. The electron configuration created is iso-electronic to a noble gas, but with a lower Z_{eff}. Adding another electron would be difficult because the (new) outer most electron has a lower attraction to the nucleus, and that electron would have to go into a higher energy shell.
This is the basis of electron , the energy needed to remove an electron in the gas phase.

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[i] This handout is based on the works of MOOG and Farrell. Moog, Richard & Farrell, John; Chemistry, A guided Inquiry; 3rd ed.;