Unit 5 – Ionic Compounds

Early Booklet E. C.:	**/ 2**
Unit 5 Hwk/Notes:	**/ 64**
Unit 5 Lab Pts.:	**/ 10**
Late, Incomplete, No work, No Units Fees? Y / N	

Learning Targets for Unit 5

1.1 I can define chemical bond
1.2 I can describe the formation of positive and negative ions
1.3 I can relate ion formation to electron configuration
1.4 I can describe the formation of ionic bonds and the structure of ionic compounds
1.5 I can generalize about the strength of ionic bonds based on the physical properties of ionic compounds
1.6 I can categorize ionic bond formation as exothermic or endothermic
1.7 I can relate a formula unit of an ionic compound to its composition
1.8 I can write formulas for ionic compounds and oxyanions
1.9 I can apply naming conventions to ionic compounds and oxyanions
1.10 I can describe a metallic bond
1.11 I can relate the electron sea model to the physical properties of metals
1.12 I can define alloys, and categorize them into two basic types.

Unit Vocabulary for Unit 5

Anion	Cation	Chemical bond	Crystal lattice
Electrolyte	Ionic bond	Ionic compound	Lattice energy
Formula unit	Monatomic Ion	Oxidation number	Oxyanion
Polyatomic ion	Binary Compound	Delocalized electron	Electron sea model
Metallic bond			

5.1 Notes:	**/ 10**
5.1 Problems:	**/ 6**
Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3	
Final Score:	**/ 16**

5.1 Problems – Ion Formation

Go to hut-lhansen.weebly.com for lesson notes.

5.1 Objectives:

1.
2.
3.
4.

1. Ion Example: Sodium.
Electrons in a sodium atom: _____.
Electron configuration: _____.
It loses one electron when ionizing, and its new electron configuration is _____.

When atoms (or group of atoms) lose or gains electrons, it becomes an _____. These are called _____ if there is only one atom involved, and _____ if more than one atom cluster carries a charge. Elements do not become other elements by losing electrons (protons don't change). _____ are positive ions, _____ are negative ions. Ionic charges are superscripted beside the symbol: **K^+, S^{2-}, Al^{3+}.**

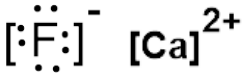
Ions follow the _____ rule, where they will tend to gain or lose valence electrons to mimic their nearest noble gas neighbor. Exceptions: _____, _____, _____ and _____, which are stable with _____ electrons.

Noble gases already have a _____ outer shell, so are unlikely to form bonds with other elements.

Cation Details: Metals lose electrons. Group 1 and 2 metals form _____ and _____ ions. Transition metals lose s-sublevel electrons, and may lose d-sublevel electrons too. The charges of different metals are written in parenthesized Roman numerals when they're named. **Ex. lead (IV) chloride.**

2. Naming Anions:
Commonly end in ide:
oxygen becomes _____,
nitrogen = _____,
phosphorus = _____.

Anion details: Nonmetals gain electrons. Group 5 & 6 elements gain 3 and 2 electrons, respectively, forming **3-** and **2-** charged ions. Group 7 (halogens) elements gain one: charge is **1-**.

Note: electron dot structure of ions shows symbol with added or lost electrons: 

3. Review! (4.4) How many valence electrons are in these elements?
cesium _____ rubidium _____
gallium _____ fluorine _____
strontium _____ oxygen _____

4. Explain why noble gases are not likely to form chemical bonds.

5. Explain why halogens and alkali metals are likely to form ions.

6. Explain how an anion of nitrogen forms, and write its electron dot structure.

Reflections. What parts of this section made sense? _____
What interested you most? _____
What parts need more explanation? _____

5.2 Notes:	**/ 10**
5.2 Problems:	**/ 6**
Late, Incomplete, No work, No Units Fee:	- 1 - 2 - 3
Final Score:	**/ 16**

5.2 Problems – Ionic Bonds

Go to hut-lhansen.weebly.com for lesson notes.

5.2 Objectives:

1.
2.
3.

1. Ionic Compound Examples
Table Salt: sodium chloride: NaCl
Fluorite: _____: _____
Lime: _____: _____

Electrostatic force that holds oppositely charged particles together is an _____ _____. A neutrally charged substance with these bonds is an _____ _____. 'Neutrally charged' means that positive and negative charges are _____.When two <u>different</u> elements combine, a _____ _____ compound forms.

Binary Compound Formula Process
1. List both ions and their charges.
2. Total positive charge must equal total negative charge: add positives or negatives until this is true.
3. Write the formula (cation first), using subscripts to show amounts of ions.

When writing chemical formulas, the _____ ion is written first. If a formula requires two or more of an ion, this number is written as a _____, a number behind and below the symbol: **Ex: K_2O**.

Ionic compounds form a _____, a repeating pattern of ions. They tend to be hard, rigid, and brittle solids, and tend to have high _____ and _____ points. They also are good _____ when dissolved in water, meaning that they conduct electricity well.

Ionic compounds can be broken apart by melting. The energy required to break apart a mole (6.02 E 23 particles) of an ionic compound is _____ _____. The amount of energy depends on the _____ and _____ of the ions present. The smaller the ion, and/or the higher the ionic charge, the greater the lattice energy.

2. Lattice Energy Example
Which has greater lattice energy, potassium fluoride (KF) or potassium iodide (KI)?
Compare ion sizes:
Potassium is used in both formulas = same size.
Fluoride is _____ than iodide, so _____ has a greater lattice energy.

3. How do ionic bonds form? In other words, what has to happen for an ionic bond to occur?

4. Determine the formulas of these binary compounds:
   A. potassium chloride, a salt substitute
   B. strontium chloride, used in fireworks
   C. sodium fluoride, used in toothpaste

5. Which binary compounds are not likely to occur: CaK, Na_2S, $BaCl_3$, MgF? Why not?

6. Which has the greater lattice energy, K_2O or CaO? Justify your choice.

<u>Reflections.</u> What parts of this section made sense? _____
What interested you most? _____
What parts need more explanation? _____

5.3 Notes:	**/ 10**
5.3 Problems:	**/ 6**
Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3	
Final Score:	**/ 16**

5.3 Problems – Names of Ionic Compounds

Go to hut-lhansen.weebly.com for lesson notes.

Start using the ions list in your resources!

5.3 Objectives:

1.
2.

The simplest ratio of the ions in a compound is a _____ _____. A negative ion with oxygen atoms connected to another element is an _____.

1. Ion Review: (5.1) Circle monatomic ions, and box the polyatomic ions:
Cl^- O^{2-} O_2^{2-} NO_3^- Mg^{2+} SO_3^{2-} Al^{3+} CrO_4^{2-} I^- Pb^{2+}
Highlight the oxyanions.

Process: Determining Ionic Formulas
Wizards vs. Monkeys!

The Wizard Method

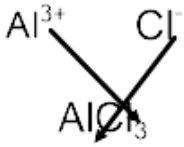
1. Write ion symbols and charges down.
2. Add positive and negative ions until net charge equals zero.

Example: aluminum chloride
aluminum ion = Al^{3+} chloride ion = Cl^-
With three chloride ions, the 3+ charge of the aluminum ion cancels. The formula is $AlCl_3$.

The Monkey Method

1. Write down the ions involved, (symbols and charges) from your ions list.
2. Swap charge values between ions, and write the formula:

Caution: The formula must be reduced if possible:
Example: Pb_2O_4 reduces to PbO_2.

aluminum chloride: 

Process: Naming Ionic Compounds From Formulas

Use your ions lists!
1. Write cation's (positive ion) name first.
2. Write the anion's name.

Note: if cation has multiple oxidation numbers, specify which one using Roman Numerals. Ex: $PbCl_2$ lead (II) chloride

2. Scavenger Hunt!

Write the names of two metals from your ions list (Resources 3) with three possible oxidation numbers (charges):

_____,_____.

NOTE: If multiple <u>polyatomic ions</u> are needed in the formula, put the ion in parenthesis, and the number of polyatomic ions written as a subscript. Ex: Ammonium Sulfate: $(NH_4)_2SO_4$

3. Write the ionic compound formulas. 2 Points
   a. calcium iodide
   b. copper (II) chloride
   c. potassium periodate
   d. silver acetate
   e. ammonium sulfate
   f. lead (II) nitrate

4. Name these ionic compounds.$CuBr_2$
   a. Mg_3N_2
   b. NaClO
   c. $Pb(SO_4)_2$

5. KNO_3There are errors in the following formulas - write the correct formulas.
   a. Na_3SO_4
   b. $BaOH_2$
   c. Fe_2O

Reflections. What parts of this section made sense? _____
What interested you most? _____
What parts need more explanation? _____

5.4 Notes:	**/ 10**
5.4 Problems:	**/ 6**
Late, Incomplete, No work, No Units Fee: - 1 - 2 - 3	
Final Score:	**/ 16**

5.4 Problems – Metallic Bonds

Go to hut-lhansen.weebly.com for lesson notes.

5.4 Objectives:

1.
2.
3.
4.

Metal atoms have _____valence electrons that move around, which is known as the 'Electron Sea Model'. As these valence electrons move around, they leave behind a charged _____ _____. The number of valence electrons a metal has equals its maximum charge. Ex: Iron has a +3 ion, so it can have 3 valence electrons. The attraction of a metal center to delocalized electrons gives rise to a _____ bond.

Metals are shiny, generally solid at room temperature, and _____, meaning that they can be flattened (due to delocalized electrons that act as lubricants during deformation). They are also _____, which means they can bend without breaking, and are _____, which means they can be drawn into wire. Finally, they transfer heat and electricity well making them _____ and _____ conductive, because their delocalized electrons are free to move when voltage is applied. Metals form _____ - mixtures of elements with metallic properties

Melting and boiling points relate to the number of valence electrons in metals: the more valence electrons, the _____ the temperature. Negative valence electrons are attracted to positive metal centers. The more electrons there are, the more of these bonds there are, so more energy is needed to break them.

1. Alloy Types:

_____ Alloy: some metal atoms replaced with atoms of similar size. Ex: Sterling silver: silver (___%) mixed with Copper (__%)

_____ Alloy: small holes in a lattice filled with small atoms. Ex: Carbon steel: carbon fits between iron atoms.

2. Draw a metal lattice of three beryllium atoms:

3. Describe how malleability and ductility of metals are explained by metallic bonding.

4. The melting point of beryllium is 1287 degrees Celsius, while that of lithium is 180 degrees Celsius. Explain the large difference in values.

5. Copper and zinc form brass, a substitutional alloy. Briefly explain why these two metals form a substitutional alloy, rather than an interstitial alloy.

6. Draw brass here:

Reflections. What parts of this section made sense? _____

What interested you most? _____

What parts need more explanation? _____

Chemistry	5.1 Lab - Build Your Own Ionic Solid!				
Name:					Correction Credit: Half
Lab Points:	Missed:	Late, No Units, No Work Fee:	First Score:	Corrections:	Final Score:
10		-1 -2 -3			

Theory:

Ionic compounds are substances made of equal numbers of positive and negative charges.

In this lab common household substances – vinegar and baking soda - will react to make one ionic and two molecular compounds according to the reaction:

$$HC_2H_3O_2 + NaHCO_3 \rightarrow NaC_2H_3O_2 + CO_2 + H_2O$$

Reactants → Products

Safety:

You will be using Bunsen burners to heat up an acidic solution. Heat glassware slowly, and pay close attention when your solution gets close to the end of its boiling.

Goggles, long pants, and close-toed shoes must be worn during the entire lab.

Clean glassware is ESSENTIAL for this lab. Even if it already looks clean, rinse out the glassware you intend to use before starting.

Equipment and Chemicals:

Ring Stand with Side Clamp
Steel Mesh
Balance and Weighing Boat
Bunsen Burner and Striker
250 mL, 100 mL Beakers
10 mL, 100 mL Graduated Cylinders
Small Test Tube and Clamp
Tongs
Glass Stirring Rod
2 mL Plastic Dropper
Boiling Stones
5% Acetic Acid (Vinegar)
Sodium Bicarbonate (Baking Soda)

Procedure:

Setup:

1. Obtain 3.0 grams of sodium bicarbonate using a pre-weighed boat.
2. Put the sodium bicarbonate into the 250 mL beaker.
3. I will deliver 50.0 mL of acetic acid into your graduated cylinder. Pour it into the 100 mL beaker. Note: you will not need all 50 mL of acid in the next part – return unused portion.

Neutralization:

4. **Slowly** and **gradually** measure and add exactly 35 mL (with graduated cylinder) of acid to your 250 mL beaker of baking soda while stirring with the rod. The liquid will foam, so pour only a little at a time. As you are stirring, you should notice the baking soda dissolving until there is no powder remaining in your reaction beaker.
5. Gently heat (do not boil yet!) your beaker with the burner. It will foam more, so stir it with the glass rod until the bubbles diminish.
6. Using the 2 mL plastic dropper, drop acid slowly into your heated solution until you see that bubbles no longer form.

Concentration:

7. Pour 3 mL of deionized water into the small graduated cylinder for later, and have a small test tube ready in your rack to collect your sample.
8. Once the reaction bubbles have stopped, add 5 boiling stones and bring the solution to a boil, stirring it the whole time.
9. When the level of liquid gets down to the top of your boiling stones, you are getting **close to the end** of your operation. Pick up the 250 mL beaker with your tongs while swirling to heat it more slowly.
10. Swirl the solution until the bottom of the beaker SUDDENLY turns white and a layer of crystals forms. Turn off the heat as soon as this happens.
11. Immediately pour the 3 mL of deionized water onto the crystals and swirl the solution around, dissolving as much crystalline residue as possible.
12. Don't wait too long, and carefully pour the hot liquid into the small test tube, leaving the boiling stones in the beaker.
13. Cool the test tube by running it under cold water (don't get water in the tube!) until it is cold to the touch.
14. Tell me when it's cool, and I will come around to test whether you followed the directions correctly by placing a seed crystal in it. If it all crystallizes, then you'll get full credit for the procedure!

Clean up:

Bring your test tubes with sodium acetate to me and I will recover and use it.
Pour the boiling stones into the waste beaker and I will wash and reuse them.
Wash the rest of your dishes and put them away.

Questions: Answer in complete sentences.

1. (2 pts) Define "polyatomic ion". You may have to use your book.

2. (3 pts) Name the three products of the reaction (see the reaction in the Theory section. You will have to use your Table of Ions Resource too).

3. (2 pts) Which of the three products has a metal and a polyatomic ion?

4. (3 pts) Describe what happens when a seed crystal is added. What happens to the temperature of the solution? Write this in at least **three** complete sentences for full credit.

Chemistry	Essential Skill 5.1 - Ionic Compounds Part 1
Name:	**Period:**
This is a practice worksheet to prepare you in the short term for the Ionic Compounds Quiz, and for the long term by reinforcing an essential skill in chemistry.	

Part 1. Use your rules of naming ionic compounds and table of common ions to write out the formulas of the following compounds:

1. calcium bromide _____
2. silver phosphate _____
3. cobalt (II) nitrate _____
4. copper (I) arsenate _____
5. nickel (II) cyanide _____
6. hydrogen peroxide _____
7. iron (III) oxide _____

Part 2. Name the following compounds. Don't forget Roman numerals for the metals with more than one oxidation state!

8. FeO _____
9. $Cd(ClO_3)_2$ _____
10. $Al(NO_3)_3$ _____
11. $KMnO_4$ _____
12. $Cu(OH)_2$ _____
13. $Sr(ClO_4)_2$ _____
14. $Co_2(HPO_4)_3$ _____

Chemistry	Essential Skill 5.2 - Ionic Compounds Part 2
Name:	**Period:**
This is a practice worksheet to prepare you in the short term for the Ionic Compounds Quiz, and for the long term by reinforcing an essential skill in chemistry.	

Part 1. Write the formulas of the following compounds:

1. ammonium oxalate _____
2. zinc bicarbonate _____
3. magnesium nitride _____
4. nickel (III) sulfate _____
5. aluminum iodide _____
6. tin (IV) dichromate _____
7. sodium chromate _____
8. lead (II) fluoride _____

Part 2. Name the following compounds. Include Roman numerals where needed.

9. $(NH_4)_2CO_3$ _____
10. $CaCl_2$ _____
11. $Pb(CrO_4)_2$ _____
12. ZnS _____
13. BaI_2 _____
14. $Be(C_2H_3O_2)_2$ _____
15. Li_2SO_4 _____
16. $NaHCO_3$ _____

Completed Points: / 10
Late/Inc. Fee: -1 -2 - 3
Final Score: / 10

Unit 5 Test Review – Ionic Compounds

This serves as test preparation for the Unit 5 exam. Points are based on completion, and we will go over any questions you have during the review.

Write formulas for the following:

1. tin (IV) chloride _____
2. lithium sulfate _____
3. aluminum oxide _____
4. aluminum cyanide _____
5. hydrogen sulfide _____
6. potassium sulfate _____
7. aluminum iodide _____
8. zinc carbonate _____

Name the following compounds.

9. $CaBr_2$ _____
10. $FeCl_2$ _____
11. CuS _____
12. $Ba(NO_3)_2$ _____
13. Cu_2Se _____
14. $Cu(OH)_2$ _____
15. $LiClO_3$ _____
16. MgF_2 _____

17. List and define four properties of metals.

18. Rank the following in order of increasing melting point: NaCl, KCl, LiF, LiCl, KBr.

19. Define, draw, and name an <u>interstitial alloy</u>.

20. Define, draw, and name a <u>substitutional alloy</u>.

21. Explain why ionic compounds shatter when hit, but metals bend.

22. Draw a metal lattice using aluminum atoms – show bonds that form between the metal centers and the delocalized valence electrons.