Name ____________________ Section ________ Date ________

Chapter 4: Basic Review Worksheet

1. When writing the name of an ionic compound, which is named first, the anion or the cation? Give an example.

2. What ending is added to the root name of an element to show that it is a simple anion in a Type I ionic compound? Give an example.

3. What general type of element is involved in Type II compounds?

4. What *two* systems are used to show the charge of the cation in a Type II ionic compound? Give examples of each system for two compounds.

5. Describe the system used to name Type III binary compounds (compounds of nonmetallic elements). Give four examples illustrating the method.

6. What is an *oxyanion*? Give two examples.

7. What is an *acid*? Give two examples (one which contains oxygen, one which does not).

8. Name each of the following binary ionic compounds.
   a. $NaCl$ c. $MgBr_2$ e. CaS
   b. K_2O d. AlI_3 f. SrO

9. Without consulting the text, name each of the following polyatomic ions.
   a. NH_4^+ c. NO_3^- e. ClO_4^-
   b. SO_3^{2-} d. OH^- f. PO_4^{3-}

10. Name each of the following compounds.
    a. NO_2 b. ICl c. CO

11. Write the formula for each of the following compounds.
    a. potassium sulfide c. calcium sulfate
    b. hydrochloric acid d. copper(II) bromide

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Chapter 4: Review Worksheet

1. How does the system used to name Type III compounds differ from that used for ionic compounds? How is the system for Type III compounds similar to that for ionic compounds?

2. What naming system is used for a series of related oxyanions to indicate the relative number of oxygen atoms in each ion? Give examples.

3. How are acids that do *not* contain oxygen named? Give three examples.

4. Describe the naming system for the oxyacids. Give an example of a series of oxyacids illustrating this system.

5. Name each of the following binary ionic compounds.
   a. $FeCl_3$ c. Al_2O_3 e. $BaCl_2$
   b. NiS d. CuO f. FeO

6. Which of the following formulas are incorrect? Correct each incorrect formula.
   a. NaS c. $NaCl_2$
   b. K_2S d. CaBr

7. Without consulting the text, name each of the following polyatomic ions.
   a. NO_2^- c. SO_4^{2-}
   b. ClO^- d. CN^-

8. Name each of the following compounds.
   a. PCl_5 b. N_2O_4 c. SF_6

9. Write the formulas for each of the following compounds.
   a. nitric acid c. silver sulfite
   b. sodium hydride d. ammonium acetate

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Chapter 4: Challenge Review Worksheet

1. What principle do we use in writing the formula of an ionic compound such as $NaCl$ or MgI_2? How do we know that *two* iodide ions are needed for each magnesium ion, whereas only one chloride ion is needed per sodium ion?

2. What is a *polyatomic* ion? Without consulting a reference, list the formulas and names of at least ten polyatomic ions. When writing the overall formula of an ionic compound involving polyatomic ions, why are parenthesis used around the formula of a polyatomic ion when more than one such ion is present? Give an example.

3. Name each of the following binary ionic compounds.
   a. Cr_2S_3 c. Fe_2O_3 e. MnO_2
   b. Cu_2S d. AuI_3 f. $CoBr_2$

4. Which of the following formulas are incorrect? Correct each incorrect formula.
   a. Rb_3N c. BaP_2
   b. Cs_3Cl_2 d. AlI_3

5. Name each of the following compounds.
   a. B_2O_3 b. P_4O_{10} c. N_2O_3

6. Write the formulas for each of the following compounds.
   a. hydrosulfuric acid c. magnesium perchlorate
   b. barium phosphate d. manganese(II) chloride

Chapter 3: Challenge Review Worksheet

1. While students certainly don't have to memorize all the elements, they should at least be able to give the symbols and names for the most common elements.

2. Dalton's atomic theory as presented in this text consists of five main postulates. Although Dalton's theory was exceptional scientific thinking for its time, some of the postulates have been modified as our scientific instruments and calculational methods have become increasingly more sophisticated. The main postulates of Dalton's theory are as follows: (1) Elements are made up of tiny particles called atoms; (2) all atoms of a given element are identical; these atoms are different from the atoms of all other elements; (4) atoms of one element can combine with atoms of another element to form a compound, and such a compound will always contain the same relative numbers and types of atoms; (5) atoms are rearranged into new groupings during an ordinary chemical reaction, and no atom is ever destroyed and no new atom is ever created during such a reaction. Students should explain these in their own words.

3. A given compound always contains exactly the same relative masses of its constituent elements. This statement is termed the law of constant composition. The law of constant composition is a result of the fact that a given compound always contains the same types and numbers of each constituent atom. For example, water's composition by mass (88.8% oxygen, 11.2% hydrogen) is a result of the fact that each water molecule contains one oxygen atom (relative mass 16.0) and two hydrogen atoms (relative mass 1.008 each). The law of constant composition is important to our study of chemistry because it means that we can always assume that any sample of a given pure substance, from whatever source, will be identical to any other sample.

4. The various isotopes of an element have virtually identical chemical properties since the chemical properties of an atom are a function of the electrons in the atom (*not* the nucleus). The physical properties of the isotopes of an element (and compounds containing those isotopes) may differ because of the difference in mass of the isotopes.

5. Isolated atoms do not form ions on their own, but are induced to gain or lose electrons by some other species (which loses or gains the electrons).

6. An ionic compound could not possibly exist of just cations or just anions: there must be a balance of charge or the compound would be very unstable (like charges repel each other).

7. a. 9p, 10n, 10e; b. 12p, 12n, 10e; c. 26p, 30n, 23e

8. a. 37p, 36e; b. 26p, 24e; c. 1p, 2e; d. 13p, 10e; e. 17p, 18e; f. 8p, 10e

9. RbH, $RbCl$, Rb_2O; FeH_2, $FeCl_2$, FeO; AlH_3, $AlCl_3$, Al_2O_3.

Chapter 4: Basic Review Worksheet

1. When naming ionic compounds, we name the positive ion (cation) first. Sodium chloride is an example.

2. For simple binary Type I ionic compounds, the ending *–ide* is added to the root name of the element that is the negative ion (anion). For example, the Type I ionic compound formed between potassium and sulfur, K_2S, is named potassium sulfide: potassium is the cation, sulfur is the anion (with the suffix *–ide* added).

3. Type II compounds involve elements that form more than one stable ion, and so it is necessary to specify which ion is present in a given compound.

4. Type II compounds are named by either of two systems, the "ous-ic" system (which is falling out of use), and the "Roman numeral" system that is preferred by most chemists. For example, iron forms two types of stable ions: Fe^{2+} and Fe^{3+}. Iron can react with oxygen to form either of two stable oxides, FeO or Fe_2O_3, depending on which cation is involved. Under the Roman numeral naming system, FeO would be named iron(II) oxide to show that it contains Fe^{2+} ions; Fe_2O_3 would be named iron(III) oxide to indicate that it contains Fe^{3+} ions. The Roman numeral used in a name corresponds to the charge of the specific ion present in the compound. Under the less-favored "ous-ic" system, for an element that forms two stable ions, the ending –ous is used to indicate the lower charged ion, so FeO and Fe_2O_3 would be named ferrous oxide and ferric oxide, respectively. The "ous-ic" system has fallen out of favor since it does not indicate the actual charge on the ion, but only that it is the lower or higher charged of the two. This can lead to confusion: for example Fe^{2+} is called ferrous ion in this system, but Cu^{2+} is called cupric ion (since there is also a Cu^{+} ion).

5. In writing the name for such compounds, the element listed first in the formula is named first (using the full name of the element), and then the second element in the formula is named as though it were an anion (with the *–ide* ending). Since there often may be more than one compound possible involving the same two nonmetallic elements, the naming system for Type III compounds goes one step further than the system for ionic compounds, by explicitly stating (by means of a numerical prefix) the number of atoms of each of the nonmetallic elements present in the molecules of the compound. For example, carbon and oxygen (both nonmetals) form two common compounds, CO and CO_2. To indicate clearly which compound is being discussed, the names of these compounds indicate explicitly the number of oxygen atoms present by using a numerical prefix.

CO	carbon *monoxide*	(*mon* or *mono* is the prefix meaning "one")
CO_2	carbon *dioxide*	(*di* is the prefix meaning "two")

The prefix *mono* is not normally used for the first element named in a compound when there is only one atom of the element present, but numerical prefixes are used for the first element if there is more than one atom of that element present. For example, nitrogen and oxygen form many binary compounds:

NO	nitrogen *mon*oxide	
NO_2	nitrogen *di*oxide	
N_2O	*di*nitrogen *mon*oxide	
N_2O_4	*di*nitrogen *tet*roxide	(*tetra* or *tetr* means "four")

6. Several families of polyatomic anions contain an atom of a given element, combined with differing numbers of oxygen atoms. Such anions are called "*oxyanions*". For example, sulfur forms two common oxyanions, SO_3^{2-} and SO_4^{2-}.

7. Acids, in general are substances that produce protons (H^+ ions) when dissolved in water. HCl and H_2SO_4 are two examples.

8. a. sodium chloride; b. potassium oxide; c. magnesium bromide; d. aluminum iodide; e. calcium sulfide; f. strontium oxide.

9. a. ammonium ion; b. sulfite ion; c. nitrate ion; d. hydroxide ion; e. perchlorate ion; f. phosphate ion.

10. a. nitrogen dioxide; b. iodine monochloride; c. carbon monoxide

11. a. K_2S; b. HCl; c. $CaSO_4$; d. $CuBr_2$

Chapter 4: Review Worksheet

1. Type III binary compounds represent compounds involving only nonmetallic elements. In writing the name for such compounds, the element listed first in the formula is named first (using the full name of the element), and then the second element in the formula is named as though it were an anion (with the *–ide* ending). This is similar, thus far, to the method used for naming ionic compounds (Type I). Since there may be more than one compound possible involving the same two nonmetallic elements, the naming system for Type III compounds goes one step further than the system for ionic compounds, by explicitly stating (by means of a numerical prefix) the number of atoms of each of the nonmetallic elements present in the molecules of the compound.

2. When there are two oxyanions in such a series (as for sulfur), the name of the anion with fewer oxygen atoms ends in *–ite* and the name of the anion with more oxygen atoms ends in *–ate*. Under this method, SO_3^{2-} is named sulf*ite* and SO_4^{2-} is named sulf*ate*. When there are more than two members of such a series, the prefixes *hypo-* and *per-* are used to indicate the members of the series with the *fewest* and *largest* number of oxygen atoms. For example, bromine forms four common oxyanions as listed.

Formula	Name
BrO^-	*hypo*bromite (fewest number of oxygens)
BrO_2^-	brom*ite*
BrO_3^-	brom*ate*
BrO_4^-	*per*bromate (largest number of oxygens)

3. For acids which do not contain oxygen, the prefix *hydro-* and the suffix *–ic* are used with the root name of the element present in the acid (for example: HCl, hydrochloric acid; H_2S, hydrosulfuric acid; HF, hydrofluoric acid).

4. The nomenclature of acids whose anions contain oxygen is more complicated. A series of prefixes and suffixes is used with the name of the non-oxygen atom in the anion of the acid. These prefixes and suffixes indicate the relative (not actual) number of oxygen atoms present in the anion. Most of the elements that form oxyanions form two such anions: for example, sulfur froms sulfite ion (SO_3^{2-}) and sulfate ion (SO_4^{2-}), and nitrogen forms nitrite ion (NO_2^-) and nitrate ion (NO_3^-). For an element that forms two oxyanions, the acid containing the anions will have the ending *–ous* if the anion is the *–ite* anion and the ending *–ic* if the anion is the *–ate* anion. For example, HNO_2 is nitr*ous* acid and HNO_3 is nitr*ic* acid; H_2SO_3 is sulfur*ous* acid and H_2SO_4 is sulfur*ic* acid. The halogen elements (Group 7) each form four oxyanions, and consequently, four oxyacids. The prefix *hypo-* is used for the oxyacid that contains fewer oxygen atoms than the *–ite* anion, and the prefix *per-* is used for the oxyacid that contains more oxygen atoms than the *–ate* anion. For example,

Acid	name	anion	anion name
HBrO	*hypo*brom*ous*	BrO^-	*hypo*brom*ite*
$HBrO_2$	brom*ous*	BrO_2^-	*hypo*brom*ite*
$HBrO_3$	brom*ic* acid	BrO_3^-	brom*ate*
$HBrO_4$	*per*bromic acid	BrO_4^-	*per*brom*ate*

5. a. iron(III) chloride; b. nickel(II) sulfide; c. aluminum oxide; d. copper(II) oxide; e. barium chloride; f. iron(II) oxide

6. a. incorrect, should be Na_2S; b. correct; c. incorrect, should be NaCl; d. incorrect, should be $CaBr_2$.

7. a. nitrite ion; b. hypochlorite ion; c. sulfate ion; d. cyanide ion

8. a. phosphorus pentachloride; b. dinitrogen tetroxide; c. sulfur hexafluoride

9. a. HNO_3; b. NaH; c. Ag_2SO_3; d. $NH_4C_2H_3O_2$

Chapter 4: Challenge Review Worksheet

1. The principle we use when writing the formula of an ionic compound is sometimes called the "principle of electroneutrality". This means that a chemical compound must have an overall net electrical charge of zero. For ionic compounds, this means that the total number of positive charges on the positive ions present must equal the total number of negative charges on the negative ions present. For example, with sodium chloride, if we realize that an individual sodium ion has a 1+ charge, and that an individual chloride ion has a 1- charge, then if we combine one of each of these ions, the compound will have an overall net charge of zero: (1+) + (1-) = 0. On the other hand for magnesium iodide, when we realize that an individual magnesium ion has a 2+ charge, then clearly one iodide ion with its 1- charge will not lead to a compound with an overall charge of zero: we would need *two* iodide ions, each with its 1- charge, to balance the 2+ charge of the magnesium ion: (2+) + 2(1-) = 0. If we consider magnesium oxide, however, we would need only one oxide ion, with its 2- charge, to balance with one magnesium with its 2+ charge [(2+) + (2-) =0], and so the formula of magnesium oxide is just MgO.

2. A polyatomic is an ion containing more than one atom. Parentheses are used in writing formulas containing polyatomic ions to indicate unambiguously how many of the polyatomic ions are present in the formula, while making certain that there is no mistake as to what is meant by the formula. For example, consider the calcium phosphate. The correct formula for this substance is $Ca_3(PO_4)_2$, which indicates that three calcium ions are combined for every two phosphate ions (check the total number of positive and negative charges to see why this is so). If we did not write the parenthesis around the formula for the phosphate ion, that is, if we had written Ca_3PO_{42}, people reading this formula might think that there were 42 oxygen atoms present.

3. a. chromium(III) sulfide; b. copper(I) sulfide; c. iron(III) oxide; d. gold(III) iodide; e. manganese(IV) oxide; f. cobalt(II) bromide

4. a. correct; b. incorrect, should be CsCl; c. incorrect, should be Ba_3P_2; d. correct

5. a. diboron trioxide; b. tetraphosphorus decoxide; c. dinitrogen trioxide

6. a. H_2S; b. $Ba_3(PO_4)_2$; c. $Mg(ClO_4)_2$; d. $MnCl_2$

Chapter 5: Basic Review Worksheet

1. Uncertainty in measurement means that we can never take an exact measurement (except for counting). The last digit of a recorded measurement is estimated, and therefore uncertain.

2. A unit tells us what scale or standard is being used to represent the results of the measurement.

3. Dimensional analysis is a method of problem solving which pays particular attention to the units of measurements and uses these units as if they were algebraic symbols that multiply, divide, and cancel. Consider the following example. A dozen of eggs costs $1.25. Suppose we want to know how much one egg costs, and also how much three dozen eggs will cost. To solve these problems, we need to make use of two equivalence statements:

   1 dozen eggs = 12 eggs

   1 dozen eggs = $1.25

   The first of these equivalence statements is obvious: everyone knows that 12 eggs are "equivalent" to one dozen. The second statement also expresses an equivalence: if you give the grocer $1.25, he or she will give you a dozen eggs. From these equivalence statements, we can construct the conversion factors we need to answer the two questions. For the first question (what does one egg cost) we can set up the calculation as follows

   $$\frac{\$1.25}{12\text{ eggs}} = \$0.104 = \$0.10$$

   as the cost of one egg. Similarly, for the second question (the cost of 3 dozen eggs), we can set up the conversion as follows