4. FORMATION OF COMPOUNDS

The noble (inert) gases – $_{2}He$, $_{10}Ne$, $_{18}Ar$, $_{36}Kr$, $_{54}Xe$ – are so called because they are chemically unreactive. In science, this is indicative of stability. The noble gases are stable so they do not need to react chemically to gain stability. We know that the reason these gases are stable is because of their electron count. You can think of the electron count of each of these elements as magic numbers: ***2***, ***10***, ***18***, ***36***, ***54***, which confers stability on the atoms.

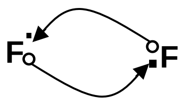
In nature, all systems tend to move to a state having the greatest stability. Compounds are formed

																	He
Li^{+}	Be^{2+}														O^{2-}	F^{-}	Ne
Na^{+}	Mg^{2+}											Al^{3+}			S^{2-}	Cl^{-}	Ar
K^{+}	Ca^{2+}											Ga^{3+}			Se^{2-}	Br^{-}	Kr
Rb^{+}	Sr^{2+}											In^{3+}			Te^{2-}	I^{-}	Xe
Cs^{+}	Ba^{2+}																Rn

4.2 Covalent Bonds

We have already said that nonmetals tend to gain e^- to attain the electron count of the nearest noble gas (to the right). This is fine when there is a nearby atom willing to supply e^-'s. If, however, no such supplier exists in the container, another method of acquiring e^-'s must be found.

Consider the case of a sample consisting only of ${}_{9}F$ atoms. ${}_{9}F$ requires one e^- to acquire the electron count of the nearest noble gas, ${}_{10}Ne$. One can think of F as having an electron hole that needs to be filled.



Each fluorine atom seeks to seize one electron from the other fluorine atom while retaining all of its own 9 electrons. The result is that the two atoms are drawn together with the two electrons, one from each fluorine atom, located between the two nuclei.

$$F \overset{\times}{\bullet} F$$

This type of bond is known as a covalent bond. In covalent bonds, the electrons are shared between two atoms.

Covalent bonds are formed between nonmetal atoms.

4.3 Nomenclature

Nomenclature is a system of naming compounds so that:

- given the name of the compound, you can write the formula of the compound uniquely; or
- given the formula of the compound, you can write the name uniquely.

Before giving the system for naming compounds, it is necessary to know something about the nature of the compound.

4.3.1 BINARY IONIC COMPOUNDS

The term ***binary compound*** means that the compound consists only of two different types of atoms.

Exercise: State whether the following compounds are binary.

4. CO
5. CO_2
6. $C_2H_2Cl_2$
7. CH_3CH_3
8. HCN

Binary ionic compounds are formed from single atom (metal) cation and single atom (non-metal) anions.

Binary ionic compound = metal + non-metal

Whenever one of the atoms is a metal the compound is an ionic compound.

e.g. Which of the following compounds are binary ionic compounds?

CO_2

CaO

$CaCO_3$

$FeCl_3$

$NaNO_2$

CCl_4

Review:
You must be able to recognize

4.3.1.1 Simple Binary Ionic Compounds

e.g. Name each of the following compounds.

metal + non-metal + **ide**

The suffix –ide is very important. It is the signature of binary compounds.

Thus there are two things to look out for:

1. Is it binary? look for the suffix ***–ide***.
2. Is it ionic? look to see if there is a metal.

Exercise: Determine whether each of the following is a binary ionic or binary covalent compound.

$NaBr$

$CaCl_2$

BaO

Li_3N

$NaNO_3$

4.3.1.2 Chemical Formulas from Names

In the previous section, we have been given the formula and from the formula we provided the name. A

$Ca \rightarrow Ca^{2+} + 2\ e^{-}$

$Cl + e^{-} \rightarrow Cl^{-}$

d. Multiply each half equation by small integers so that the number of electrons produced in the 1st equation equals the number of electrons absorbed in the 2nd equation

Each Cl absorbs only one e^{-}. Thus we need 2 Cl's to absorb both electrons produced by the Ca.

$Ca \rightarrow Ca^{2+} + 2\ e^{-}$

$2 \times [Cl + e^{-} \rightarrow Cl^{-}]$

e. Put together the formula

$CaCl_2$

Exercise: Write the formula for each of the following compounds.

magnesium chloride

barium sulphide

aluminium selenide

rubidium oxide

strontium fluoride

calcium bromide

indium oxide

e.g., Write the name of each of the following compounds.

1. Cr_2O_3
   a. Determine the oxidation state of each element in the compound
      - Use Rule 1 to set up an equation

$$2(Cr) + 3(O) = 0$$

      - This is an equation with two unknowns. You must choose a value of oxidation number for one of the atoms. The priority is given to Group A atoms. Choose oxygen according to its group number.

$$(O) = -2$$

      - Substitute this into the equation above.

$$2(Cr) + 3(-2) = 0$$
$$2(Cr) = +6$$
$$(Cr) = +3$$

   b. Write the name of the compound.
      It is a binary, ionic compound.
      Cr has a charge of +3

chromium (III) oxide

2. $CoBr_6$
   a. Determine the oxidation state of each element in the compound
      - Use Rule 1 to set up an equation

$$2(Co) + 6(Br) = 0$$

      - This is an equation with two unknowns. You must choose a value of oxidation number for one of the atoms. The priority is given to Group A atoms. Choose bromine according to its group number.

$$(Br) = -1$$

      - Substitute this into the equation above.

$$(Co) + 6(-1) = 0$$
$$(Co) = +6$$

   b. Write the name of the compound.
      It is a binary, ionic compound.
      Co has a charge of +6

cobalt (VI) bromide

3. SnS_2
   a. Determine the oxidation state of each element in the compound
      - Use Rule 1 to set up an equation

$$(Sn) + 2(S) = 0$$

      - This is an equation with two unknowns. You must choose a value of oxidation number for one of the atoms. The priority is given to Group A atoms. Choose sulphur according to its group number.

$$(S) = -2$$

      - Substitute this into the equation above.

$$(Sn) + 2(-2) = 0$$
$$(Sn) = +4$$

   b. Write the name of the compound.
      It is a binary, ionic compound.
      Sn has a charge of +4

tin (IV) sulphide

e.g., Write the formula for each of the following compounds

1. Iron (III) chloride
   a. Identify the charges of the ions

   $$(Fe) = +3 \qquad (Cl) = -1$$

   b. Combine the two types of atoms so there is a net charge of zero

   $$Fe \rightarrow Fe^{3+} + 3\ e^{-}$$

   $$3 \times [\ Cl + e^{-} \rightarrow Cl^{-}\ ]$$

   c. Write the formula: $FeCl_3$

2. Manganese (V) oxide
   a. Identify the charges of the ions

   $$(Mn) = +5 \qquad (O) = -2$$

   b. Combine the two types of atoms so there is a net charge of zero

   $$2 \times [\ Mn \rightarrow Mn^{5+} + 5\ e^{-}\ ]$$

   $$5 \times [\ O + 2\ e^{-} \rightarrow O^{2-}\ ]$$

   c. Write the formula: Mn_2O_5

3. Mercury (II) nitride
   a. Identify the charges of the ions

   $$(Hg) = +2 \qquad (N) = -3$$

   b. Combine the two types of atoms so there is a net charge of zero

   $$3 \times [\ Hg \rightarrow Hg^{2+} + 2\ e^{-}\ ]$$

   $$2 \times [\ N + 3\ e^{-} \rightarrow N^{3-}\ ]$$

   c. Write the formula: Hg_3N_2

****

General Rule for the Nomenclature of all Binary Ionic Compounds

e.g., Write the names of each of the following compounds

K_2S

W_2O_5

CdO

$ZnCl_2$

AgCl

4.3.3 BINARY COVALENT COMPOUNDS

These can be recognised because they contain no metal atoms. For our purposes, only nonmetals form covalent bonds.

Binary covalent compounds:

1. NO

POLYATOMIC IONS

Ammonium

Mercury (I)

Permanganate

Dichromate

Chromate

Sulfate

Sulfite

Phosphate

Phosphite

Nitrate

Nitrite

Acetate

Oxalate

Hydroxide

Cyanide

Thiocyanate

Perchlorate

Chlorate

Chlorite

Hypochlorite

Carbonate

4.3.4 Naming Acids

Acids (Arrhenius definition) – substances which, when dissolved in water, produce H^+ ions.

By implication then, acids are ionic compounds in which the cations are all protons, H^+.

Nomenclature of Acids is based on the nomenclature of the anion as there is no need to reference the cation which is always H^+.

Anion	Corresponding Acid	Example	
		Anion	Corresponding Acid
—ide	hydro—ic acid	__	

4.3.5 Acid anions

For anions having multiple charges, there is no reason for the cations to be identical. It is perfectly legitimate to have an ionic compound with a formula that combines several different cations. For example, the combination of lithium and ammonium cations with phosphate can have:

Exercises on Nomenclature

For each substance whose name is given, write the formula, if the formula is given, provide the name. Use the Periodic Table as your main reference. If you encounter ions with which you are unfamiliar, you may look these up in your textbook.

This assignment will count as part of the quiz grade. It is due on Thursday, March 18, 2004.

1. carbon dioxide ________
2. $BaCl_2$ ________
3. calcium hypochlorite ________
4. $KMnO_4$ ________
5. barium hydroxide ________
6. $FePO_4$ ________
7. cobalt (II) acetate ________
8. SnS ________
9. calcium fluoride ________
10. CBr_4 ________
11. mercury (II) chlorate ________
12. $Al_2(Cr_2O_7)_3$ ________
13. Ammonia ________
14. $Al_2(SO_4)_3$ ________
15. manganese (II) oxalate ________
16. $CdCl_2$ ________
17. sodium carbonate ________
18. KH_2PO_4 ________
19. mercury (I) chloride ________
20. $Al(ClO_2)_3$ ________
21. aluminium oxide ________
22. $Zn(NO_3)_2$ ________
23. gallium selenate ________
24. $CaCl_2 \cdot 6H_2O$ ________
25. iron (III) sulphide ________
26. NH_4NO_2 ________
27. iron (II) sulphate heptahydrate ________
28. H_3PO_3 *(aq)* ________
29. barium bromate ________
30. H_2O_2 ________
31. lead (II) iodide ________
32. $Ca(OH)_2$ ________
33. carbonic acid ________
34. K
35. sodium sulphite ________
37. periodic acid ________
39. magnesium hydrogen sulphate ________