* Notes + practice key

Ionic Compounds

Situation #1 – Metal (only 1 ion charge) + Non-metal

- Name the metal off the periodic table Name the non-metal changing the ending to "ide" *Example* – $MgCl_2$ *magnesium chloride*
- To write the formula, balance the ion charges and write the simplest ratio of + to – ions as subscripts. (do not include charges in formulas!)

Example – aluminum sulfide

Ions: Al^{3+} and S^{2-} so… we need 2 Al^{3+} and 3 S^{2-} (+6 balances -6)
Formula = Al_2S_3

Situation #2 – Multivalent Metal + non-metal

- Most transition metals are multivalent (more than one possible ion charge). Zn^{2+} and Ag^{+} are exceptions having one possible charge.
- Name the metal off the periodic table including a **roman numeral** to show which ion charge is used.
- Name the non-metal, changing the ending to "ide"

Example: $FeCl_2$ – iron (II) chloride
(We have two Cl^- ions (total charge -2) and only 1 Fe ion, so charge must be +2 to balance -2)

Example: Copper (I) Oxide
Ions: Cu^{+1} *and* O^{2-} *… so we need 2* Cu^{+1} *ions to balance 1* O^{2-} *ion*
Formula: Cu_2O

Situation #3 – Polyatomic ions (one or both ions in an ionic compound are polyatomic)

- Ammonium (NH_4^+) can replace a metal and there are many negative polyatomic ions (ex. SO_4^{2-}) can replace a non-metal ion in an ionic compound.
- Name them exactly as written on the chart. Do not change ending to "ide"
- If there is more than 1 of a polyatomic ion needed to balance charges, put it in **brackets!**

Example: Calcium Nitrate → $Ca(NO_3)_2$
Ca^{2+} NO_3^-

Situation #4 – Ionic Hydrates

- Ionic compounds can form weak bonds with water molecules to form crystals
- The number of water molecules bonded is indicated with a prefix followed by "hydrate"
- The weak bond is indicated with a dot

Example: Copper (II) sulphate pentahydrate
$CuSO_4 \cdot 5H_2O$

1. Determine the correct formula for each of the following **ionic compounds**. *Some are hydrates!*

a. rubidium sulphide ..... Rb_2S

b. manganese (II) chloride ..... $MnCl_2$

c. sodium oxalate ..... $Na_2C_2O_4$

d. ammonium dichromate ..... $(NH_4)_2Cr_2O_7$

e. zinc nitrate tetrahydrate ..... $Zn(NO_3)_2 \cdot 4H_2O$

f. tungsten (VI) sulphate ..... $W(SO_4)_3$

g. copper (I) monohydrogen phosphate ..... Cu_2HPO_4

h. manganese (IV) bisulphate ..... $Mn(HSO_4)_4$

i. calcium acetate ..... $Ca(CH_3COO)_2$

j. iron (III) carbonate hexahydrate ..... $Fe_2(CO_3)_3 \cdot 6H_2O$

k. tin (IV) oxide ..... SnO_2

l. ammonium dihydrogen phosphate ..... $NH_4H_2PO_4$

m. calcium sulphite octahydrate ..... $CaSO_3 \cdot 8H_2O$

n. nickel (III) chlorate trihydrate ..... $Ni(ClO_3)_3 \cdot 3H_2O$

2. Determine the correct name for each of the following **ionic compounds**. *Some are hydrates!*

a. Rb_3PO_4 ..... rubidium phosphate

b. $Fe(OH)_3$ ..... iron (III) hydroxide

c. $Cr_2(SO_4)_3 \cdot 5H_2O$ ..... chromium (III) sulfate pentahydra

d. $NH_4HC_2O_4$ ..... Amonium binoxalate

e. $Pb(ClO)_4$ ..... Lead (IV) hypochlorite

f. $Ba(OH)_2 \cdot 10H_2O$ ..... Barium hydroxide decahydrate

g. $Hg_2(NO_3)_2$ ..... Mercury (I) nitrate

h. Cu_2SO_3 ..... Copper (I) sulphite

i. $Ni_2(Cr_2O_7)_3$ ..... Nickel (III) dichromate

j. $Ta(ClO_4)_2$ ..... Tantalum (II) perchlorate

Acids vs. Bases

Acids	Bases
Aqueous H^+ compounds	Aqueous OH^- compounds
Turn blue litmus red	Turn red litmus blue
H^+ appears first in the formula (expect acetic acid CH_3COOH)	OH^- appears last in the formula

Naming Acids

- ***If the negative ion (anion) ends in "ide"*** → hydro_____ ic acid.
- Example: HCl
  IUPAC name = Hydrogen Chloride
  Classic name = Hydrochloric Acid

- ***If the negative ion (anion) ends in "ate"*** → _____ ic acid
- Example: H_2SO_4
  IUPAC name = Hydrogen Sulphate
  Classic name = Sulphuric acid

- ***If the negative ion (anion) ends in "ite"*** → _____ous acid
- Example: $HClO_2$
  IUPAC name = Hydrogen Chlorite
  Classic acid name = Chlorous Acid

Name the following:

a. H_2SO_3 (sulphite) sulpherous acid
b. HCN (cyanide) hydrocyanic acid
c. HF (fluoride) hydrofluoric acid
d. HNO_3 (Nitrate) nitric acide

Naming Bases

- Named as normal ionic compound
- Will always end in hydroxide (unless they are organic bases which we will worry about in our organics unit!)
- Example: NaOH → Sodium Hydroxide

Covalent Compounds

- Non-metal atoms share electrons forming a covalent bond which holds them together in a group called a molecule
- Molecular formulas show the # of each type of atom in a molecule

Molecular Elements

- Some elements always form molecules

H_2 N_2 O_2 F_2 Cl_2 Br_2 I_2 P_4 S_8

Molecular Compounds

- You are ALWAYS told how many of each atom through the prefix naming system. (mono, di, tri, tetra, penta, hexa, hepta, octa, nona, deca)
- Covalent compounds DO NOT form ions so you do not need to worry about balancing ion charges!!
- Please note: mono is not used on the first element, only the second. (example: Carbon monoxide)

Examples: dinitrogen oxide = N_2O
sulfur dioxide = SO_2
Nitrogen = N_2
Sucrose = $C_{12}H_{22}O_{11}$

Write the formulas for each covalent molecule in the reactions below. Include phase subscripts:
solid = **(s)** liquid = **(l)** gas = **(g)** aqueous (dissolved in water) =**(aq)**

1. Solid silicon reacts with gaseous fluorine to produce gaseous silicon tetrafluoride.

$Si(s) + F_2(g) \rightarrow SiF_4(g)$

2. Hydrogen and oxygen gases react to produce dihydrogen monoxide.

$H_2(g) + O_2(g) \rightarrow H_2O(l)$

3. Solid sucrose combusts in the presence of oxygen gas to produce carbon dioxide gas and water vapour.

$C_{12}H_{22}O_{11}(s) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

4. Dihydrogen monoxide diluted in water decomposes in the presence of a catalyst the produce oxygen gas and water.

$H_2O_2(aq) \rightarrow O_2(g) + H_2O(g)$