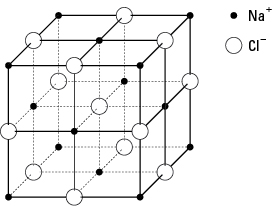
Chemistry Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Mrs. Pavlovich Mon/Thurs Period\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Unit 6 – Chemical Names and Reactions**Kavanah Topic 2 pages 25 – 44

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1. Chemical Symbols
   1. Agreed upon by International Union of Pure and Applied Chemists \_\_\_\_\_\_\_\_\_\_\_
   2. Most elements are written without a subscript except for diatomic elements.
      1. Always exist in nature as 2 atoms together.
      2. Use a subscript 2.
      3. BrINClHOF
2. Compounds
   1. Combinations of elements ex. H2O, SO3, ~~N~~~~2~~(N2 is not a compound. It is an element.)
   2. Chemically combined in definite proportions by mass
   3. Formulas indicate which elements are present and the number of atoms of each.
   4. Subscript is written to the lower right of element
   5. Subscript tells you the number of that type of atom in the compound
   6. If no subscript, assume that there is only one of that atom of that element   
      ex. H2SO4 has one sulfur atom
   7. Practice: How many of each type of atom are in the molecule H2SO4?  
      Hydrogen – 2 atoms  
      Sulfur – 1 atom  
      Oxygen – 4 atoms
   8. Practice: How many of each type of atom are in the compound (NH4)2S?  
      nitrogen – 2 atoms  
      hydrogen – 8 atoms  
      sulfur – 1 atom
3. Empirical formulas
   1. Simplest integer ratio of atoms in the compound, either covalent or ionic
   2. Ionic compounds
      1. Metal and non-metal
      2. Electrons are transferred
      3. Form a crystal of ions, not molecules
      4. Ionic compounds are always represented with empirical formulas  
           
           
         Homework: Kavanah p. 30 # 2-11. Bring Kavanah book to class tomorrow and every day.
   3. Molecules
      1. Have covalent bonds
      2. Non-metals
      3. Electrons are shared
      4. Can be reduced to an empirical formula.
         1. Ex. What is the molecular formula of butane? C4H10
         2. What is the empirical formula of butane? C2H5
      5. A molecular formula may be a multiple of an empirical formula
4. Ions
   1. Remember that atoms and compounds are electrically neutral
   2. Ions are charged particles
   3. Cations attract anions in a ratio that produces a neutral compound
   4. The superscript to the right of the symbol gives the charge
      1. Write the number first ex. Al3+ not Al+3
      2. If there is no number but just a plus or minus sign, then the charge is 1
   5. Monatomic ions
      1. When one atom gains or loses an electron to form an ion
      2. Top oxidation state is usually the charge of the ion
      3. Sn, Pb, and many transition elements form several different possible ions
      4. Names of anions end in –ide  
         Ex. What is the name of the ion formed from oxygen?
   6. Polyatomic ions
      1. A group of atoms covalently bonded together, possession a charge
      2. Find these polyatomic ions in Table E and write the formulas:  
         sulfate \_\_\_\_\_\_\_\_\_  
         hydronium \_\_\_\_\_\_  
         carbonate \_\_\_\_\_\_\_
      3. Most of the polyatomic ions are negative
      4. Parentheses are used to enclose polyatomic ions when there is more than one
      5. Subscript after parentheses tells how many of that polyatomic ion
      6. Practice: Name the polyatomic ions in (NH4)3PO4? How many of each?
      7. Practice: How many of each type of *atom* are in the molecule (NH4)3PO4?
   7. Ionic Compounds
      1. Formed from monatomic and/or polyatomic ions
      2. Ratio of ions produces a neutral compound ex. NaOH
      3. Ionic compounds are not molecules
   8. Coefficients
      1. Sometimes written in front of formula
      2. Applies to the entire formula
      3. Tells how many units of the formula are present
      4. Ex. 3NaOH
         1. 3 formula units of NaOH (don’t use the word molecule)
         2. 3 sodium atoms
         3. 3 oxygen atoms
         4. 3 hydrogen atoms
      5. Practice: How many hydrogen atoms are in 2(NH4)3PO4?
   9. Hydrates of Ionic Compounds
      1. Crystal lattice of ionic compounds binds water into the structure
      2. Hydrates have a definite number of water molecules for each unit of the compound
      3. Ex. barium chloride incorporates 2 water molecules into crystal  
         BaCl2·2H2O
      4. Anhydrous compound (without water) formed by heating. Write equation for dehydration of barium chloride.  
           
           
         Homework: Kavanah p. 30 1-11 (multiple choice) AND 14 and 15.
5. Formulas and Names of Ionic Compounds
   1. Equalize charges to write formulas
      1. Compounds must have equal number of + and – charges
      2. Criss-cross method: write the absolute value of the ion charge as the subscript of the other ion. Use parentheses for polyatomic ions if needed  
         Ex. magnesium ion and chloride ion  
         Ex. ammonium ion and sulfide ion
      3. Reduce to empirical formula  
         Ex. aluminum ion and nitride ion

Homework: Ion POGIL

* 1. The stock system
     1. Used for metals with more than one oxidation state
     2. Sn, Pb, and many of the transition elements
     3. Use a Roman numeral to indicate the charge of the metal ion
     4. Ex. What is the charge of Fe ion? \_\_\_\_\_\_ Co ion? \_\_\_\_\_
     5. Ex. Iron III oxide  
        Ex. copper II sulfate
  2. Naming Binary Ionic Compounds
     1. Binary compounds contain 2 elements
     2. Write the cation (usually the metal) first
     3. The anion (ending in –ide) comes second
     4. Ignore the subscript numbers
     5. Ex. write the names:  
        MgO  
        MgCl2
     6. If compound contains a transition element or Sn or Pb, reverse criss-cross to find the charge of the ion, then Roman numeral.
     7. Practice: Find the charge of the metal ion and then write the name:  
        FeCl2  
        FeCl3NiO  
        Cr2O3
     8. Do NOT use Roman numeral if there is only one oxidation state. Ex:  
        AgCl  
        Zn2S

1. Four types of reactions
   1. Chemical Equations  
      ex. C(s) + O2(g) 🡪 CO2 (g)  
      1. Written form of a chemical reaction
      2. Similar to algebraic equation, because one side “equals” the other side
      3. A plus sign separates each compound
      4. The arrow is read “produces” or “yields”
      5. Reactants are on the left
      6. Products are on the right
      7. Atoms are the same on both sides, but rearranged
      8. The phase of each substance is often indicated  
         g – gas  
         s – solid  
         l – liquid  
         aq – aqueous, dissolved ionic compound
   2. Double Replacement Reactions  
      ex. AgNO3 (aq) + NaCl(aq) 🡪 AgCl(s) + NaNO3(aq)  
      1. 2 reactants are soluble ionic compounds
      2. 2 products
         1. 1 is usually a soluble ionic compound
         2. the other must be a precipitate, a gas, or a molecular compound such as water
      3. Dating analogy – 2 couples go to dance together and change partners
      4. Ex. all the reactions from the double replacement lab
      5. If both products are still in the aqueous phase, then there is NO REACTION
      6. Use Table F to see if either of the products are insoluble (solid precipitate)
      7. There is a reaction if one product is molecular, ex. H2O or H2S
   3. Synthesis (combination) Reactions  
      ex. 2H2(g) + O2(g) 🡪 2H2O(g)
      1. 2 or more reactants
      2. 1 product
      3. ex. iron rusting  
         4Fe(s) + 3O2(g) 🡪 2Fe2O3(s)
      4. Reactants can be elements or compounds  
         ex. CO2(g) + H2O(l) 🡪 H2CO3(aq)
      5. General form of synthesis is A + B 🡪 AB

* 1. Decomposition Reactions
     1. Opposite of a synthesis reaction
     2. One reactant
     3. 2 or more products
     4. Practice: Write the 3 examples above in reverse
     5. General form of decomposition is AB 🡪 A + B
  2. Single Replacement Reactions
     1. One element replaces another element in a compound
     2. Dating analogy – a couple and a single guy go to a dance…
     3. Example copper replaces silver ions in silver nitrate  
          
        Cu(s) + 2AgNO3(aq) 🡪 Cu(NO3)2(aq) + 2Ag(s)
     4. Look at Table J – Cu is above Ag. Therefore copper is more active than silver. Copper can replace silver.
  3. Naming ionic compounds with polyatomic ions
     1. Write the cation first, write the anion second
     2. Write the name of the polyatomic ion exactly as it appears in Table E
     3. Ignore the subscript numbers
     4. Ex. write the names:  
        NaNO3  
        (NH4)3N
  4. Naming binary covalent compounds
     1. This is a molecule, not an ionic substance
     2. Formed from 2 nonmetals
     3. The name of the compound ends in –ide
     4. Use prefixes to indicate how many of each atom are in the compound
     5. Common prefixes for covalent naming (not in reference tables)  
        1 mono-  
        2 di-  
        3 tri-  
        4 tetra-  
        5 penta-  
        6 hexa-  
        7 hepta-  
        8 octo-  
          
        ex. N2O4 is dinitrogen tetroxide
     6. If the element name starts with a vowel (ex. oxygen) then the final a or o of the prefix is not used
     7. Exception: do not use mono- if there is only one of the first element.   
        ex. CO carbon monoxide  
        ex. SO3 sulfur trioxide  
          
        homework: Kavanah p. 32-33 multiple choice

1. Chemical Reactions and Equations
   1. Review chemical change vs. physical change
      1. Ice melting P
      2. water boiling P
      3. water splitting into hydrogen and oxygen C
      4. salt dissolving in water P
      5. a piece of sodium is cut in pieces P
      6. a piece of sodium reacts with water C
      7. gasoline burns C
   2. What makes a chemical change?
      1. The substances produced are quite different than the starting materials
      2. The change cannot be easily reversed
      3. Look for a color change, heat, fizzing, bubbles, or smell
      4. a well-defined chemical change is a chemical reaction
   3. Balancing Chemical Equations
      1. Law of conservation of mass
         1. Matter is neither created nor destroyed in a chemical reaction
         2. The number of each type of atom is the same on both sides of the equation
      2. Some equations are balanced as soon as you write them. What now! See example above.
      3. NEVER change subscripts of compounds to balance an equation
      4. Change the coefficients to balance the equation.
      5. If no coefficient is written, the coefficient is one.
      6. List each element on the reactant side. List in the same order on the product side.
      7. Tally the number of each type of atom on each side.
      8. Change the coefficients until the numbers are equal. Do this in PENCIL
      9. Lots of examples:  
           
         ex. C(s) + O2(g) 🡪 CO (g)  
           
         ex. H2(g) + O2(g) 🡪 H2O(g)  
           
         ex. CH4 + O2 🡪 CO2 + H2O
   4. If a polyatomic ion stays the same on both sides of the reaction, balance it as a unit.  
        
      ex. Ba(OH)2 + K2CrO4 🡪 BaCrO4 + 2KOH  
        
      ex. sodium phosphate + lead II nitrate 🡪 sodium nitrate + lead II nitrate
   5. Endothermic and Exothermic Processes will not be on this test