Structure of the Atom

Formulas & Theorems Covered Today:

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Homework:

- Complete all of the components below for the first 20 elements by making a landscape page
- Look for any trends that may appear

Notes:

Atomic Diagrams

- All matter is composed of atoms and these were thought to be the smallest particles that a substance is could be broken down into until the discovery of nuclear reactions.
- Atoms are composed of even smaller particles called subatomic particles.

Subatomic particles

Particle Name	Shorthand	Charge	Mass	Location
Protons	P	+	1 AMU	Nucleus
Electrons	e-	-	1/1836 that of the proton	Atomic Orbitals
Neutrons	N	Neutral	1 AMU	Nucleus

Standard Atomic Notation

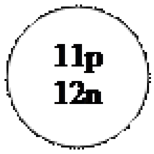
- This is a way to express an element's mass and atomic # in a simple compact way.
- It also can give us a great deal of information about an elements and its atomic structure.
- The mass is written to the left of the element's symbol as a superscript and the atomic number is written to the left as a subscript.

Ex Cobalt Iodine Nickel

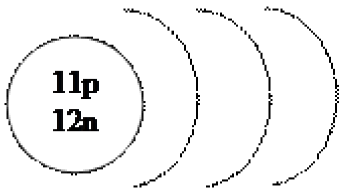
Bohr-Rutherford Diagrams

- The Bohr-Rutherford Diagrams are a combinations of the atomic theories of Neils Bohr and Ernest Rutherford.
- They pictorially represent the atomic structure of the elements.
- These diagrams have several steps that are required in order to complete them.
- We will use the element sodium as the example element but as long as the steps are followed this will work for every elements.

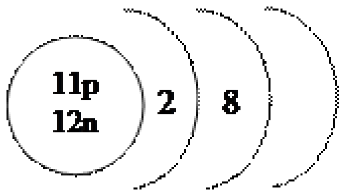
- Step 1: Add the number of protons electrons and neutrons into a circle to represent the nucleus. The information can be gathered from the standard atomic notation above



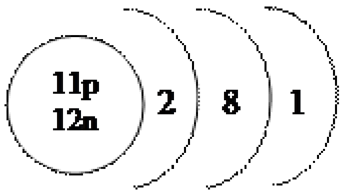
- Step 2: The next step is to determine the number of rings around the atoms that will represent the orbitals in which the electrons will exist at. These also represent the energy levels that the electrons can exist at. The electrons can only exist on these orbitals because of the work of Neils Bohr which determined that electrons had fixed amounts of energy (called quanta) and could only exist at specific distances from the nucleus of an atom. To determine the number of energy levels around an atom, you need to determine which period of the periodic table that element is situated. In the case of sodium, we know that it is found in energy level 3.



- Step 3: The next step is to fill the inner orbitals with the electrons. The inner orbitals will be filled first as stated by the aufbau principle (look up on the internet for clarification). To determine the number of electrons that each orbital can hold, you can look at the number of elements in each of the periods. You will notice that the first period of the table has only two elements, which also represents the maximum number of electrons that can fit in the first orbital. The second period has eight elements and thus the second ring can hold eight electrons and so on for all of the atomic energy levels.



- Step 4: The last step is to determine the number of electrons in the outer shell (valence shell). To do this, you need to count the number of elements in from the left that the element of interest is in that particular row. For sodium, it is the first element in the row, so the number of valence electrons that it has is one. If we were doing this for oxygen as a second example, it is number six in that row, and therefore it would have six valence electrons.



Ionic Bohr-Rutherford Diagrams

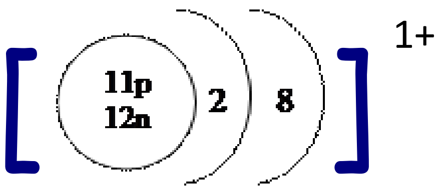
- Atoms, with the exception of the noble gases, are not stable unless they are bonded to other elements.
- To do this, many of them form ions.
- The ultimate goal of the atoms is to achieve noble gas configuration, or in other words, they want to fill their outer shells.
- When you complete the outer shell an atom is said to have a complete octet.
- When the Bohr-Rutherford Diagram has been completed, you must ask yourself the following question.

"Is it easier for ______________ to lose ______valence electrons or gain ________ valence electrons?"

- If we were to fill out the question for sodium, you would have the following.

"Is it easier for sodium to lose one valence electrons or gain seven valence electrons?"

- To answer the question, we have to remember that to lose an electron or to gain an electron takes the same amount of energy.
- Atoms will always choose the path of least resistance and so they will always chose the easier method.
- In this case it is easier for sodium to lose the one electron to achieve its noble gas configuration than it is for it to gain seven to also achieve the noble gas configuration.
- Once this question has been answered you can complete the change in the diagram as shown below.



- Notice that the diagram has been put in between square brackets and that a charge has been associated with the ion that was formed.
- When wanting to determine the charge associated with an ion, you have to compare the number of protons in the nucleus to the number of electrons in the orbitals around the atoms.
- If you recall, protons have a plus one charge and electrons have a minus one charge.
- By determining the difference in the number of protons and electrons, you can then determine the charge on the atom.

- In the case of sodium, there are eleven protons in the nucleus and ten electrons in the orbitals around the nucleus.
- You can determine that there is one more proton than electrons and therefore there is a plus one charge.

Ionic Bonding and Lewis Structures

Formulas & Theorems Covered Today:

Homework:

Ionic Bonding

- Ionic bonding occurs when two ions of opposite charges are attracted to one another through electrostatic attractions.
- Ions are formed when atoms gain and lose electrons to become isoelectric to Nobel gases.
- In order to do that, atoms that tend do lose electrons will donate their electrons to other atoms that tend to gain electrons
- The atom that has lost an electron gains a net positive charge and becomes a cation while the atom that loses an electron gains a net negative charge and becomes an anion.

Valence Electrons

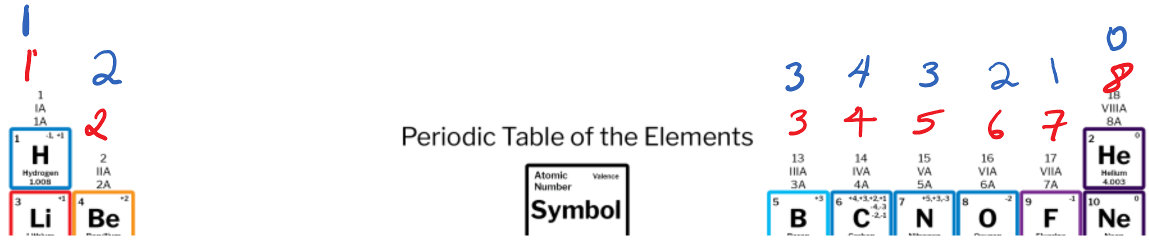
- As you learned last year and at the end of the last lesson, electrons that occupy the outer most orbital (shell) are called the valence electrons and subsequently, the outer orbital is called the valence shell
- These are of particular interest to chemists because they are the electrons that are involved, in one way or another, in chemical bonding.
- They also produce the chemical properties that you may have learned about last year.
- In order to be more efficient in our understanding of how atoms bond, and to be able to display how bonding occurs in a diagram, we have to simplify our Bohr-Rutherford diagrams into Lewis Structures
- Lewis Structures are named after Gilbert N. Lewis who, in a 1916 article called "*The Atom and the Molecule*" introduced these simplified structures to show how atoms bond together.
- When representing Lewis Structures of single elements, the symbol of the element is placed on the page and the equivalent number of valence electrons are added to the symbol.
- The electrons are added first at "12 o'clock" and then at 3, 6 and 9 o'clock in that order. (Note that if there are only 2 electrons that it would stop at 3 o'clock, and 3 electrons at 6 o'clock etc.)
- Once you have added the 4th electron, any subsequent electrons are added in the same order starting at 12 o'clock, pairing up with the electron that is already in that position

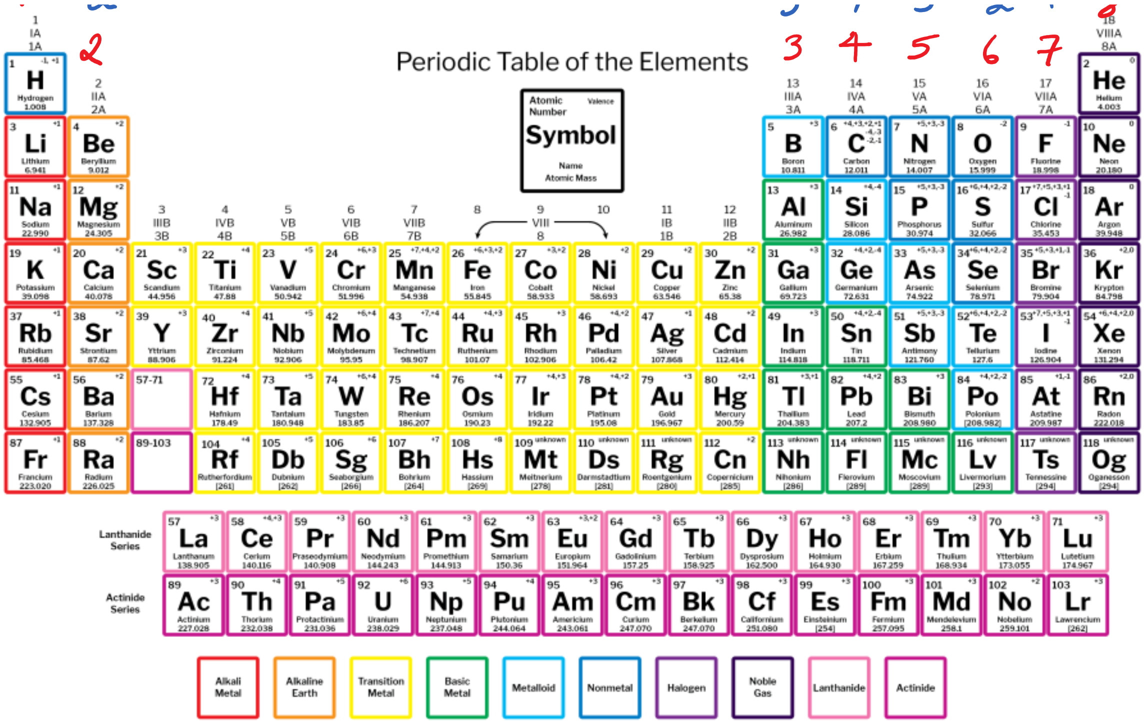
Ex.

Na	Mg	Al	Si	P	S	Cl	Ar
Na	Mg	Al	Si	P	S	Cl	Ar

valence # — 1

valence e^- — 1 2

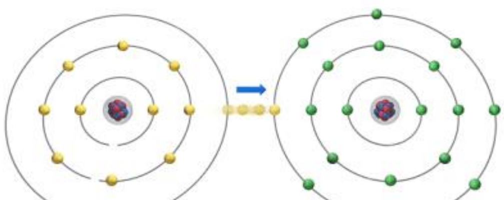


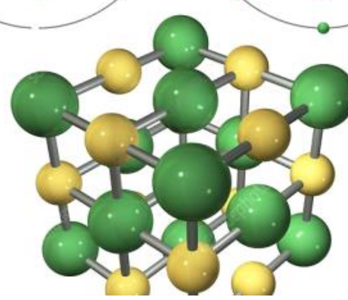


- After completing the Lewis Diagram for the elements, you will notice that every atom on the periodic table with the exception of the Nobel gases will have electrons in the structure that are unpaired.
- These unpaired electrons represent the valence number that is assigned to that element
- A valence numbers is another way of saying that this is the number of ionic bonds that that particular element can make
- For example, if we look at magnesium atoms, they have 2 valence electrons, both of which are unpaired resulting in a valence number of 2
- Similarly sulfur atoms have 6 valence electrons but only have 2 electrons that are unpaired electrons resulting in a valence number of 2

Lewis Structures of Ionic Compounds

- When creating Ionic Lewis structures, you have to complete them in 2 steps
- In the first step, you must show the transfer of electrons (where they are coming from and where they are going to)
- In the second step, you must depict the ionic compound as it appears in ion form
- Remember that ionic compounds are made of charged particles that form some type of chemical array as is depicted to the right





- Ex.

Na + Cl

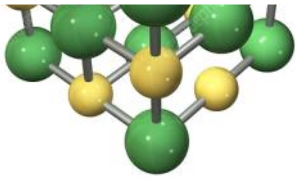
Na· ·Cl

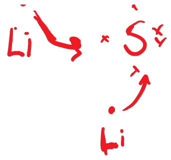
$[Na]^+ [:Cl:]^-$ = 0

Li + S

Li· ·S

Li + S





$[Li]^{+} [S]^{2-} [Li]^{+} = 0$

Mg + P

$[P]^{3-} [Mg]^{2+} [Mg]^{2+} [P]^{3-} [Mg]^{2+} = 0$

Ionic Compound Nomenclature

Formulas & Theorems Covered Today:

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- ★
- ★

Homework:

- Complete Associated Naming and Formulas Document

Notes:

Chemical Compounds

- There are two main types of chemical compounds, molecular and ionic.
- Molecular compounds are made of atoms which are tightly bound together.
- Generally, molecular compounds are made from two or more nonmetallic atoms.
- Ionic compounds are made of positive and negatively charged ions, which are held together with electrostatic attractions.
- Cations, or positive ions, are formed when metal atoms lose electrons.
- Anions, or negative ions, are formed when nonmetal atoms gain electrons.
- Ionic compounds are generally made from metallic plus nonmetallic ions.

Monatomic Ions

- The ionic charges of monatomic ions often can be determined by using the periodic table.
- The formulas of monatomic ions are written as the element symbol followed by the charge written as a superscript.
- Charges are written as the numerical value followed by the + or – sign.
- If the numerical value of the charge is one, only the + or – is written.
- Metals in Groups 1A, 2A, and 3A lose electrons when they form cations.
- Their ionic charge is positive and numerically equal to the group number.
- Their names are the same as the metal name (as in the sodium ion, Na+).
- Group A nonmetals form anions, and their charge can be obtained by subtracting 8 from the group number; the sign is negative.
- Their names end in –ide (as in oxide ion, O_2^-). Nonmetals in Group 4A and Group O normally do not form ions.

Naming Rules for Binary Monovalent Compounds

1. Name the metal (cation) in the chemical as its name appears on the periodic table
2. Identify the non-metal (anion) from the periodic table
3. Change the non-metal's ending to "ide"

$NaBr$ — Sodium Bromide

$MgCl_2$ — Magnesium Chloride

Sr_3P_2 — Strontium Phosphide

Rb_2S — Rubidium Sulphide

AlF_3 — Aluminum Fluoride

Creating Binary Monovalent Formulas

1. Write down the symbols for the elements involved leaving a space between the elements
2. In the inner upper corners of the elements, place the valence numbers for the elements
3. Criss cross the numbers
4. Reduce if possible

Lithium Fluoride	Barium Iodide	Calcium Oxide	Beryllium Phosphide
$Li^1\ F^1$	$Ba^2\ I^1$	$Ca^2\ O^2$	$Be^2\ P^3$

Lithium Fluoride | Barium Iodide | Calcium Oxide | Beryllium Phosphide

$Li^1 \times F^1$ | $Ba^2 \times I^1$ | $Ca^2 \times O^2$ | $Be^2 \times P^3$

LiF | BaI_2 | CaO | Be_3P_2

Polyatomic Ions

- Polyatomic ions are tightly bound groups of atoms that behave as a unit and carry a charge.
- The ammonium ion is a polyatomic cation.
- The names of most polyatomic anions end in either –ite or –ate.
- There are a few exceptions, including cyanide (CN^-) and hydroxide (OH^-).
- If polyatomic anions containing oxygen exist as an –ite/-ate pair, the charge on the pair is the same, and the –ite ending indicates one less oxygen atom than the –ate ending.
- Nitrite is NO_2^- and nitrate is NO_3^- .
- For a series of polyatomic anions containing oxygen (oxyanions) containing more than two members, the ion with the largest number of oxygen atoms has the prefix per- and the suffix –ate; the ion with the smallest number of oxygen atoms has the prefix hypo and the suffix -ite.
- The oxyanions containing chlorine are the most common examples:

ClO_4^- is the perchlorate ion
ClO_3^- is the chlorate ion
ClO_2^- is the chlorite ion
ClO^- is the hypochlorite ion

Naming Rules for Ionic Compounds containing Polyatomic Ions

1. Name the cation (metal or ammonium) as the name appears on the periodic table or the table of polyatomic ions
2. Name the anion (non-metal or polyatomic ion) as the name appears on the periodic table or the table of polyatomic ions (make sure to account for any name changes for the oxyanions)

$NaClO_3$ – Sodium Chlorate
$MgBrO_2$ – Magnesium bromite
$(NH_4)_2SO_2$ – Ammonium hyposulphite
$Ca_3(AsO_4)_2$ – Calcium arsenate

Creating formulas for Ionic Compounds containing Polyatomic Ions

1. Write down the symbols for the elements or polyatomic ions involved leaving a space between the elements
2. In the inner upper corners of the elements or polyatomic ions, place the valence numbers for the elements (note: that the polyatomic ions valence number can be found next to the ions on the polyatomic ion sheet)
3. Criss cross the numbers (note: if the number being placed in the subscript of the polyatomic ion is large than 1, you must place the polyatomic ion in brackets)
4. Reduce if possible

Lithium Cyanide – $Li^1 \times CN^1 = LiCN$
ammonium dichromate – $NH_4^1 \times Cr_2O_7^2 = (NH_4)_2Cr_2O_7$
aluminum perphosphate – $Al^3 \times PO_5^3 = AlPO_5$
beryllium fluorite – $Be^2 \times FO_2^1 = Be(FO_2)_2$

Multivalent Metals

- Many of the transition metals (Group B) form cations with more than one charge.
- This is also a characteristic of the cations of tin and lead, the two metals of Group 4A. (Exceptions: Ag forms only 1^+ cations; Zn and Cd form only 2^+ cations).
- There are two methods of naming transition metal ions having a variable charge.
- In the Stock system of naming, a Roman numeral in parentheses is used in the ion's name to indicate the numerical value of the charge.
- The Fe^{2+} ion is named the iron(II) ion. (There is no space between the element name and the parenthesis.)
- In the classical system of naming, the classical name of the element (such as ferr-, from ferrium, Latin for iron) is used as a root word, and a suffix is used to indicate the charge.
- The root word is followed by the suffix –ous to name the cation with the lower of the two ionic charges, and the suffix –ic is used with the higher of the two ionic charges.
- The Fe^{2+} ion is named the ferrous ion and the Fe^{3+} ion is named the ferric ion.

Naming Rules for Ionic Compounds Containing Multivalent Metals

1. Name the metal as the name appears on the periodic table or the table of polyatomic ions
2. Name the anion (non-metal or polyatomic ion) as the name appears on the periodic table or the table of polyatomic ions (make sure to account for any name changes for the oxyanions)
3. To determine the roman numeral for the metal
   a. uncross the subscripts
   b. Compare the value given to the anion to that on the periodic table or the polyatomic ion sheet.
      i. If they match use the number given to the metal as the roman numeral and place it between the name for the metal and non-metal
      ii. If they do not match, use a multiplier to equate the value to what it is on the periodic table or polyatomic ion sheet, do the same for the subscript given to the metal and then use that as the roman numeral for the metal placing it between the metal and non-metal

1 ✓ 2 ✓
Cu_2O_1 – copper(I) oxide

1×2 1×2
CuO_1 – copper(II) oxide

2×2 1×2
PbO_2 lead(IV) oxide

5 2 ✓
$Mn_2(SO_3)_5$ – manganese(V) sulphite

1×3 1×3
ScP_1 – Scandium(III) phosphide

3 ✓ 1 ✓
$Au(FO_3)_3$ – gold(III) flurrate

Creating formulas for Ionic Compounds containing Multivalent Metals

1. Write down the symbol for the multivalent metal and the anion (element or polyatomic ion) leaving a space between them.
2. In the inner upper corners of the metal and anions, place the valence numbers for the elements (note: that the multivalent metal's valence number can be found in the roman numeral of the name)
3. Criss cross the numbers (note: if the number being placed in the subscript of the polyatomic ion is large than 1, you must place the polyatomic ion in brackets)
4. Reduce if possible

gold(III) chloride – $Au^3\ Cl^1$ – $AuCl_3$

Chromium(VI) phosphite – $Cr^6\ PO_3^3$ – $Cr(PO_3)_2$

copper(II) sulphate – $Cu^2\ SO_4^2$ – $CuSO_4$

Molecular Bonding and Molecular Nomenclature

Formulas & Theorems Covered Today:

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Homework:

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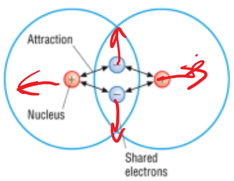
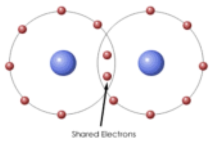
Notes:

Molecular Bonding

- In a covalent bond you have 2 elements, both of which do not want to give up any electrons
- Because both atoms want to hold onto their electrons, the atoms move closer to one another where their valence shells begin to overlap
- This is why another name for molecular bonds is co-valent bonds
- These bonds form a delicate balance between the attractive forces between electrons and nuclei and the repulsive forces between both electrons and nuclei to each other
- In order to satisfy the octets of the atoms, both atom overlap their valence shells and produce an area of increased electron density
- As a result of this, both atoms are able to use the overlapped area and the electrons within that area to complete their octets.
- Note that under normal circumstances, for every electron that one atom donates, another atom has to donate one as well
- Atoms can also make multiple bonds with other atoms that electrons are shared with but cannot exceed 3 bonds

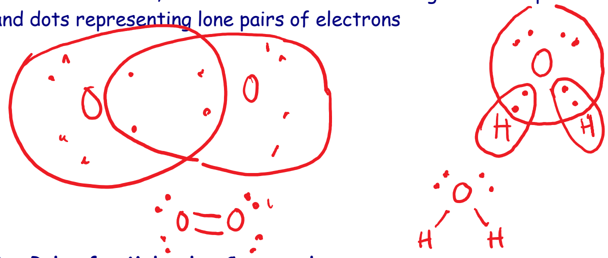


Example of a Covalent Bond

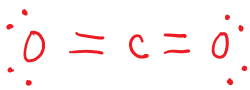


Lewis Structures of Molecular Compounds

- When wanting to create molecular Lewis structures, determine the number of electrons that need to be shared, draw out the structures using lines to represent pairs of electrons shared and dots representing lone pairs of electrons



CO_2



N_2

:N≡N:

Naming Rules for Molecular Compounds

- molecular compounds contain ONLY NON-METALS (this includes hydrogen)
- electrons are shared between the elements in the compound therefore there are NO CHARGES on these elements
- to name a molecular compound, always write the name of the element that is less electronegative (farthest to the left on the periodic table) first and the element
- that is more electronegative (closer to the right on the periodic table) second the end of the second element changes to –ide (i.e. carbon dioxide)
- if there is more then one atom of an element then you use prefixes to identify the element quantity
- Prefixes are NOT used for IONIC COMPOUNDS!!!

CO_2 – carbon dioxide
CO – carbon monoxide
C_2H_4 – dicarbon tetrahydride
SF_6 – sulfur hexafluoride
P_2O_5 – diphosphorus pentaoxide
P_4O_{10} – tetraphosphorus decaoxide

Creating formulas for Ionic Compounds containing Multivalent Metals

P_2O_5 diphosphorus p

P_4O_{10} – tetraphosphorus decaoxide

Creating formulas for Ionic Compounds containing Multivalent Metals

- use the prefix indicated on each element to determine the quantity of each element
- remember "mono-" is not always included because it represents one

nitrogen dioxide – NO_2

carbontetrachloride – CCl_4

Chemical Reactions - Word Equations

Formulas & Theorems Covered Today:

Homework:

Notes:

Chemical Reactions

- Chemical Change: is the transformation of one or more substances into new substances, with new properties.
    - 5 Signs of a Chemical Reaction
        - Colour Change
        - Heat or Light is Given Off
        - Gas/Bubbles are formed
        - A precipitate forms
        - Difficult or impossible to reverse

- Chemical Reaction: is a process by which chemical change occurs. Some reactions will absorb energy "endothermic" and others will release energy "exothermic" in the form of heat, light and/or sound.

Writing Equations

- Starting substances in a chemical reaction are called REACTANTS.
- The new substances formed in a chemical reaction are called PRODUCTS.
- The arrow in the equation means "produces" or "YIELDS".
- Sometimes heat or energy is needed or formed in a reaction.
- Heat or energy is included in the word equation just like any other reactant or product.
- Behind the name of each reactant and product you give the state of the substance in brackets using the short form indicated is also provided in an equation. For example: solid (s), liquid (l), gas (g), and aqueous (aq). *Remember: aqueous means that the substance was dissolved in water or that it is in solution form.
- A chemical reaction is often described by writing a chemical equation, word equation or balanced chemical equation.

Chemical Equations

- uses either word or symbols and formulas to describe the changes occurring during a chemical reaction. It is a summary of reactions between reactants and products.
- For example: reaction b/w solid magnesium metal and hydrochloric acid

    - Word Equation: uses names of compounds & elements
        - Magnesium + Hydrochloric acid yields Magnesium chloride + Hydrogen gas

- Skeleton Equation: uses basic formulas and symbols and is not balanced
    - $Mg_{\square} + HCl_{\square} \rightarrow MgCl_2 + H_2$

- Balanced Chemical Equation: use formulas and symbols, but is balanced
    1. includes states of each substance using symbols in brackets
    2. coefficients to balance the chemical equations
        - $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

Word Equations

Write the following word equations into skeleton chemical equations:

1. zinc + lead (II) nitrate yield zinc nitrate + lead

   $Zn + Pb(NO_3)_2 \rightarrow Zn(NO_3)_2 + Pb$

2. aluminum bromide + chlorine yield aluminum chloride + bromine

   $AlBr_3 + Cl_2 \rightarrow AlCl_3 + Br$

3. sodium phosphate + calcium chloride yield calcium phosphate + sodium chloride

   $Na_3PO_4 + CaCl_2 \rightarrow Ca_3(PO_4)_2 + NaCl$

4. potassium chlorate when heated yields potassium chloride + oxygen gas

   $KClO_3 \xrightarrow{\Delta} KCl + O_2$

5. hydrogen gas + oxygen gas yields water

   $H_2 + O_2 \rightarrow H_2O$

6. aluminum + hydrochloric acid yields aluminum chloride + hydrogen gas

   $Al + HCl \rightarrow AlCl_3 + H_2$

7. calcium hydroxide + phosphoric acid yields calcium phosphate + water

   $Ca(OH)_2 + H_3PO_4 \rightarrow Ca_3(PO_4)_2 + H_2O$

8. copper + sulfuric acid yields copper (II) sulfate + water + sulfur dioxide

   $Cu + H_2SO_4 \rightarrow CuSO_4 + H_2O + SO_2$

9. hydrogen + nitrogen monoxide yields water + nitrogen gas

   $H_2 + NO \rightarrow H_2O + N_2$

10. silver nitrate + magnesium chloride yields silver chloride + magnesium nitrate

    $AgNO_3 + MgCl_2 \rightarrow AgCl + Mg(NO_3)_2$

11. iron (III) chloride + sodium hydroxide yields iron (III) hydroxide + sodium chloride

    $FeCl_3 + NaOH \rightarrow Fe(OH)_3 + NaCl$

12. nitrogen gas + hydrogen gas yields ammonia

    $N_2 + H_2 \rightarrow NH_3$

13. phosphorus + oxygen gas yields diphosphorus pentoxide

$N_2 + H_2 \rightarrow NH_3$

13. phosphorus + oxygen gas yields diphosphorus pentoxide

$P + O_2 \rightarrow P_2O_5$

14. potassium + magnesium bromide yields potassium bromide + magnesium

$K + MgBr_2 \rightarrow KBr + Mg$

15. hydrochloric acid + calcium carbonate yields calcium chloride + water + carbon dioxide gas

$HCl + CaCO_3 \rightarrow H_2O + CO_2 + CaCl_2$

Six Types of Chemical Reactions

Formulas & Theorems Covered Today:

Homework:

Notes:

Intro to Types of Chemical Reactions

In order for substances to undergo chemical changes, they need to be reacted together in chemical reactions. These are some main types/patterns of common chemical reactions:

Decomposition: One substance breaks apart into 2 or more substances.

Generic Equation	$AB \rightarrow A + B$
Example Chemical Equation	$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

Synthesis: Two reactants combine to make a new substance.

Generic Equation	$A + B \rightarrow AB$
Example Chemical Equation	$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_\ell$

Single Displacement: One element takes the place (displaces) of another element in a compound. Like stealing a dance partner.

Generic Equation	$A + BC \rightarrow AC + B$ (when A is a cation) $A + BC \rightarrow BA + C$ (when A is an anion)
Example Chemical Equation	$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$

Double Displacement: Cations of two compounds exchange places to form 2 new compounds. Like switching dance partners.

Generic Equation	$AB + CD \rightarrow AD + CB$
Example Chemical Equation	$Pb(NO_3)_{2(aq)} + KI_{(aq)} \rightarrow PbI_{2(s)} + KNO_{3(aq)}$

Combustion Reaction: This is the rapid combination of a substance (fuel) with oxygen involving the production of heat and light. Specifically the combustion of a hydrocarbon deals with the rapid combination of oxygen with a compound containing at least carbon and hydrogen.

Generic Equation	$hydrocarbon + oxygen\ gas \rightarrow carbon\ dioxide + water$
Example Chemical Equation	$C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

Neutralization Reaction: The combination of an alkaline substance with an acidic substance to produce water and the salt of the alkaline and acidic substances.

Generic Equation	$Acid + Base \rightarrow Salt + water$
Example Chemical Equation	$HCl_{aq} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)}$

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