Unit 3: Compounds and Chemical Reactions

Chapter 6: Chemical Names and Formulas

6.1: Introduction to Chemical Bonding

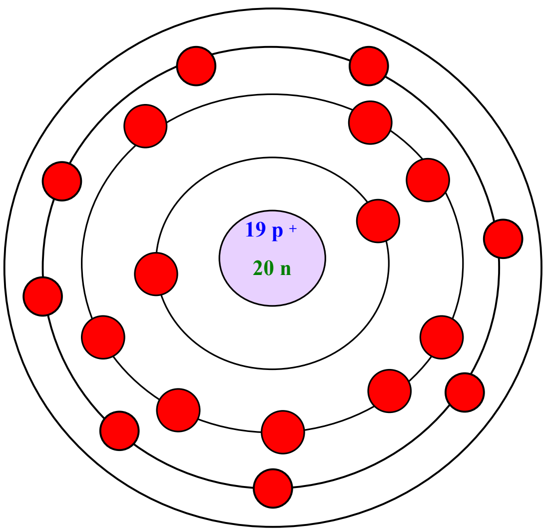
Molecules: - basic unit of a compound.
- contain at least two atoms of the same or different kind of elements.

Ions: - when atoms lose or gain electrons, they attain a positive or negative charge.

1. **Cations**: - **positive** charged ions (atoms that lose electrons).
   - naming cation (element name follow by "ion")

Example 1: Draw the energy level diagrams for the following cations.

a. Potassium ion = K^+ (19 p^+ and 18 e^-)



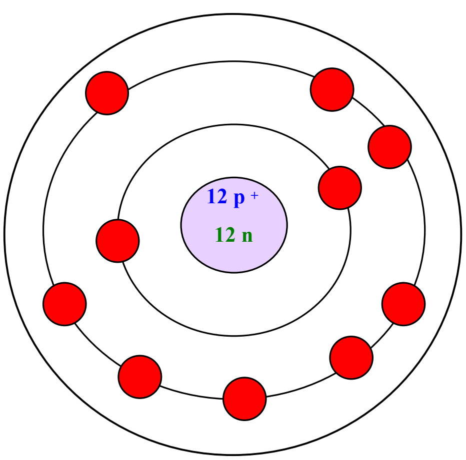
Potassium Ion (K^+)

Atomic Number: 19 **Atomic Mass: 39.10** **Nucleus: 19 p^+ and 20 n**

3rd Energy Level: 8 e^- (8 valence e^- - Filled)
2nd Energy Level: 8 e^-
1st Energy Level: 2 e^-
Total: 18 e^- **Net Charge = 1+**

Location on the Period Table of Elements: Fourth Row; Column IA

b. Magnesium ion = Mg^{2+} (12 p^+ and 10 e^-)



Magnesium Ion (Mg^{2+})

Atomic Number: 12 **Atomic Mass: 24.31** **Nucleus: 12 p^+ and 12 n**

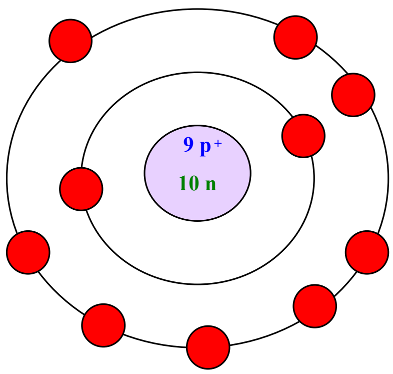
2nd Energy Level: 8 e^- (8 valence e^- - Filled)
1st Energy Level: 2 e^-
Total: 18 e^- **Net Charge = 2+**

Location on the Period Table of Elements: Third Row; Column IIA

2. **Anions**: - **negative** charged ions (atoms that gain electrons).
   - naming anion (first part of element name follow by suffix ~*ide*)

Example 2: Draw the energy level diagrams for the following anions.

a. Fluoride = F^- (9 p^+ and 10 e^-)



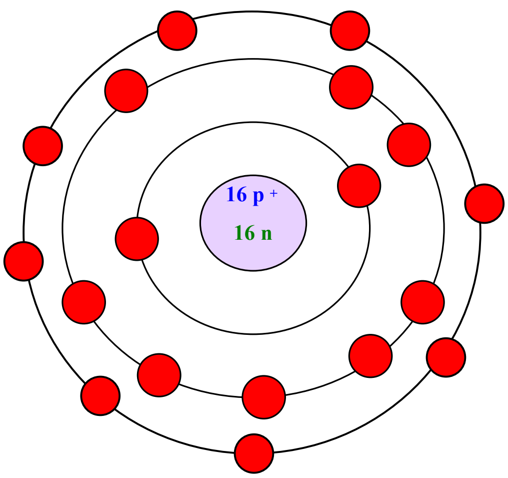
Fluoride (F^-)

Atomic Number: 9 **Atomic Mass: 19.00** **Nucleus: 9 p^+ and 10 n**

2^{nd} Energy Level: 8 e^- (8 valence e^- - Filled)
1^{st} Energy Level: 2 e^-
Total: 10 e^- **Net Charge = 1–**

Location on the Period Table of Elements: Second Row; Column VIIA

b. Sulfide = S^{2-} (16 p^+ and 18 e^-)



Sulfide (S^{2-})

Atomic Number: 16 **Atomic Mass: 32.01** **Nucleus: 16 p^+ and 16 n**

3^{rd} Energy Level: 8 e^- (8 valence e^- - Filled)
2^{nd} Energy Level: 8 e^-
1^{st} Energy Level: 2 e^-
Total: 18 e^- **Net Charge = 2–**

Location on the Period Table of Elements: Third Row; Column VIA

Octet Rule: - the tendency for electrons to fill the second and third energy levels (8 valence electrons – for s and p orbitals) to achieve stability.
- in most cases, this means having the same electron configuration of the **nearest** noble gas.
- exceptions to the rule include helium (only 2 electrons to fill the first energy level), and the transition metals.
- we only look at the s and p orbitals as valence electrons for the representative elements in higher energy levels beyond the third row.

Example: Bromide (Br^-) – (35 p^+ and 36 e^-)

Atomic Number: 35 **Atomic Mass: 32.01** **Nucleus: 35 p^+ and 16 n**
Location on the Period Table of Elements: Fourth Row; Column VIIA

4^{th} Energy Level: 18 e^- (8 valence e^- in s and p orbitals – Filled – 10 e^- in d orbitals)
3^{rd} Energy Level: 8 e^-
2^{nd} Energy Level: 8 e^-
1^{st} Energy Level: 2 e^-
Total: 36 e^- **Net Charge = 1–**

6.2: Representing Chemical Compounds

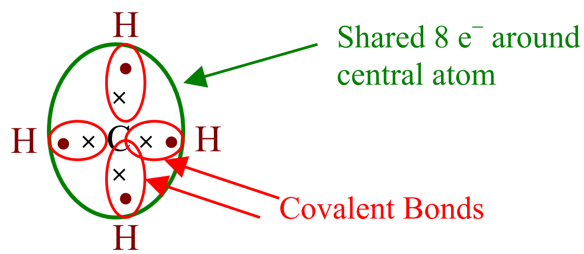
Chemical Formula: - displays the different atoms and their numbers in the smallest representative unit of a substance (includes all molecular compound, monoatomic, binary and polyatomic elements).

Molecular Formula: - chemical formula that represents the actual molecule of a molecular compound.

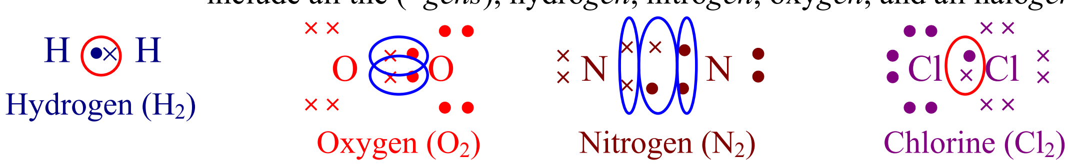
Formula Units: - chemical formulas that represent the basic unit of an ionic compound (compounds that made up of a metal and a non-metal)

Molecular Compound: - when a non-metal combines with a non-metal.
- forms **covalent bonds** (electrons are "share" between atoms).
- forms **non-electrolytes** (do not dissociate into ions when dissolve in water)

Example: CH_4



Binary Elements: - non-metals that come in pairs of atoms as molecules to achieve stability.
- include all the (~*gens*), hydro*gen*, nitro*gen*, oxy*gen*, and all halo*gen*s.



Polyatomic Elements: - non-metals that comes in groups of 4 or 8 to achieve stability.
- **P_4 (Phosphorus)** and **S_8 (Sulphur)**

Diatomic Elements (all the ~gens, including Halogens - second last column of the Periodic Table**)**	**Polyatomic Elements**	**Monoatomic Elements**
Hydrogen (H_2), Oxygen (O_2), Nitrogen (N_2), Fluorine (F_2), Chlorine (Cl_2), Bromine (Br_2), Iodine (I_2)	Phosphorus (P_4) Sulphur (S_8)	All other Elements. **Examples**: Helium (He), Iron (Fe), Calcium (Ca), Silver (Ag), Mercury (Hg)

Note: **Students should memorize all the diatomic and polyatomic elements. They are the only exceptions. All other elements are monoatomic.** A lot of the symbols are recognisable from the name of the elements (Zinc: Zn; Carbon: C; Aluminium: Al). Some of them look somewhat different. This is because the symbols came from the elements' Latin names (Silver: Ag for "Argentum"; Gold: Au for "Aurum"). To save time, students should also familiarise themselves with the whereabouts of the elements on the Table.

Law of Definite Proportion: - the same compound always contains exactly the same proportion of elements by mass. (*Proust*)

Example 1: Water contains about 8 parts oxygen to 1 part hydrogen by mass. A 192 g of unknown liquid composed of hydrogen and oxygen contains 12 g of hydrogen. Is the unknown liquid water? Justify your response.

192 g total − 12 g of hydrogen = 180 g of oxygen

$$\frac{180\text{ g oxygen}}{12\text{ g hydrogen}} = \frac{15\text{ parts oxygen}}{1\text{ part hydrogen}}$$

Since the ratio between oxygen and hydrogen is 8:1 in water, the unknown liquid is NOT water.

Law of Multiple Proportion: - when two elements form a series of compounds, the ratios of the masses of the second element that combine with the first element can always be reduced to small whole numbers. (*Dalton*)

Example 2: State the ratios of hydrogen between the following hydrocarbon compounds.

Hydrocarbons	Mass of Hydrogen per 1 g of Carbon
Compound A	2.973 g
Compound B	3.963 g
Compound C	4.459 g

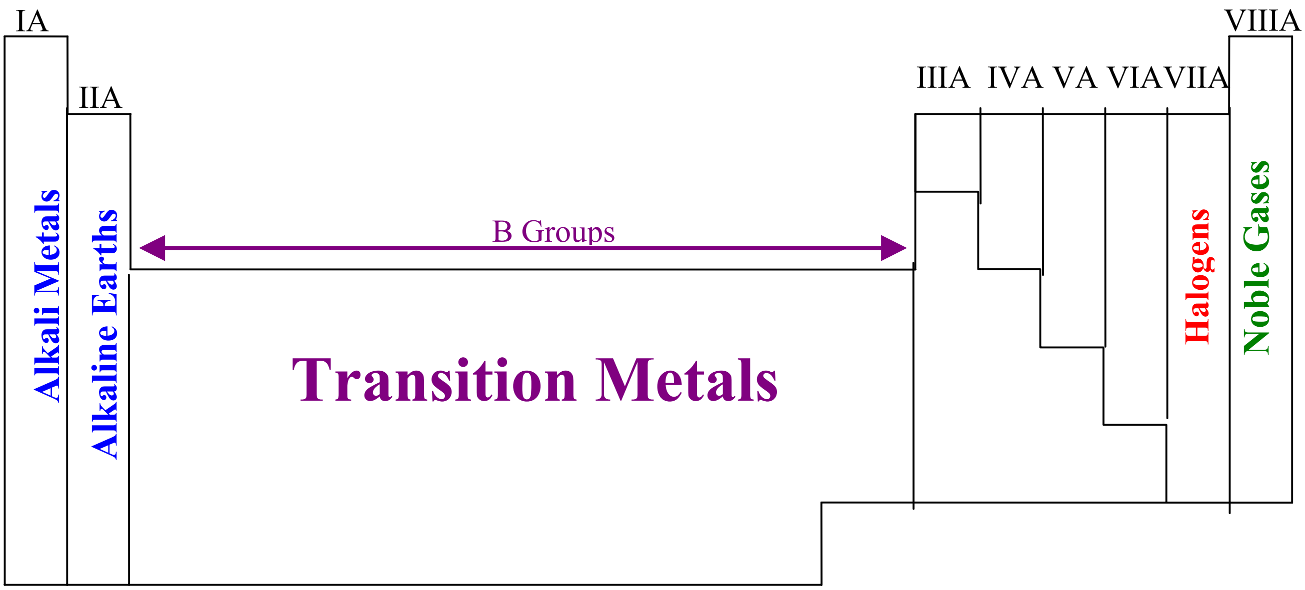
$$\frac{A}{B} = \frac{2.973\text{ g}}{3.963\text{ g}} = 0.7501892506 \approx 0.75 \qquad \frac{A}{B} = \frac{3}{4}$$

$$\frac{B}{C} = \frac{3.963\text{ g}}{4.459\text{ g}} = 0.8887642969 \approx 0.888\ldots \qquad \frac{B}{C} = \frac{8}{9}$$

$$\frac{C}{A} = \frac{4.459\text{ g}}{2.973\text{ g}} = 1.49983182 \approx 1.5 \qquad \frac{C}{A} = \frac{3}{2}$$

Assignment
6.1 pg. 137 #1 to 9
6.2 pg. 142 #10, 11, 14 and 15

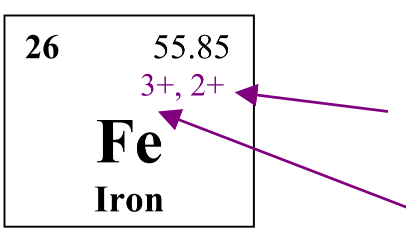
6.3: Ionic Charges



Groups or Families	Chemical Properties
Alkali Metals (IA)	very reactive; forms ions with +1 charge when react with non-metals
Alkaline Earth Metals (IIA)	less reactive than alkali metals; forms ions with +2 charge when react with non-metals
Halogens (VIIA)	very reactive; form ions with –1 charge when react with metals; all form diatomic molecules
Noble Gases (VIIIA)	very stable; do not form ions; monoatomic gas at room temperature

Transition Metals (1B to 10B): - groups and periods of metals that can have **varying charges**.
- use **Roman Numerals** as part of their ionic names.

Example: Fe^{3+} and Fe^{2+} ions



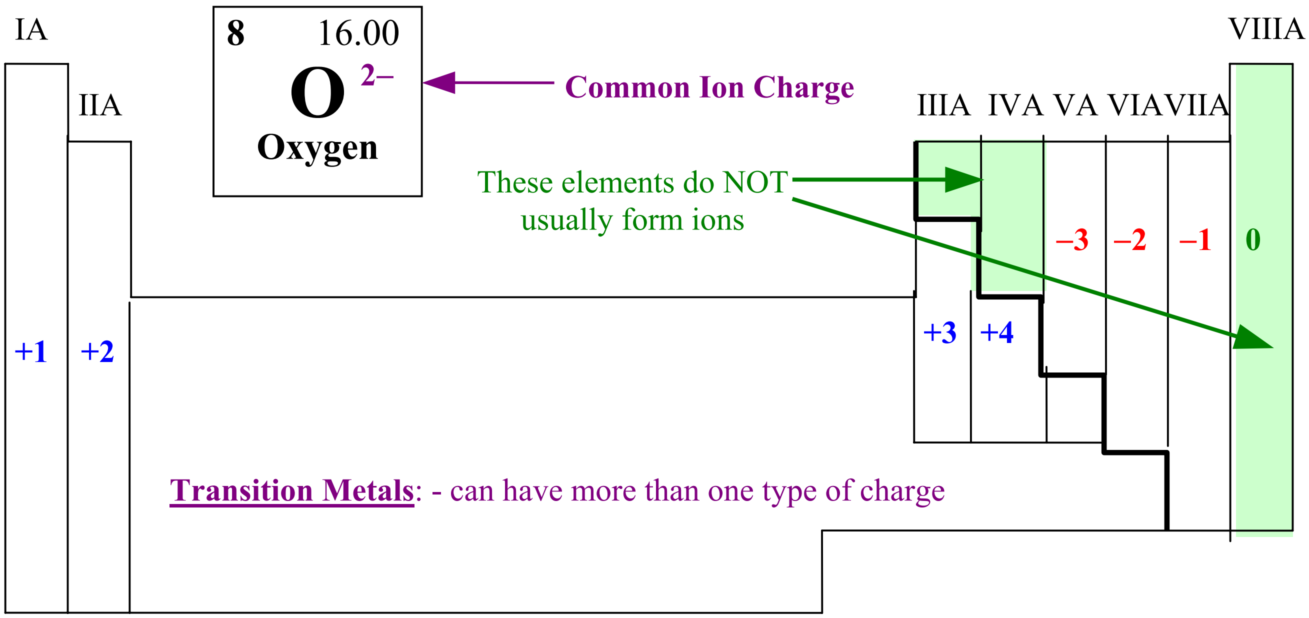
Fe^{2+} = Iron (II) ion

Fe^{3+} = Iron (III) ion
(First Charge Listed is the most common type)

Chemical Properties of Metals and Non-Metals:

1. **Metals lose electrons to become positive ions – cations**.
2. **Non-Metals gain electrons to become negative ions – anions**.
3. Hydrogen usually loses an electron to become a H^+ ion. However, it can sometimes gain an electron to become H^- (Hydride).
4. **The last column of the Table of Elements does not usually form ions**. These elements are called the **Noble Gases** (Helium, Neon, Argon, Krypton, Xenon, and Radon).

The number of electrons an atom loses or gains depends on which column (vertical) the element is at the Table.



The reason that noble gases (column VIIIA) do not form ions is because their outermost shells are filled with the maximum number of electrons allowed. That is why we call this group of elements "noble gases". They do not form ions because they are stable. Hence we use the word "noble" to describe them. All the other elements form ions because they want to achieve stability like the noble gases. If you observe carefully, oxide has the same number of electrons as the nearest noble gas, neon. On the other hand, calcium ion has the same number of electrons as the nearest noble gas, argon. In terms of stability, which is another word for lower energy state, these ions are more stable than their respective atoms.

Since the number of valence electrons of an atom is the same as its column number, all the elements of column IA have 1 valence electron. As we see with lithium, all they have to do is to lose that valence electron to achieve a noble gas "like" state. For elements in column IIA, they all have 2 valence electrons. Hence, they lose 2 electrons to acquire stability and become ions with a net charge of +2. The following table summarises these points.

Column	Number of Valence Electrons	Methods to achieve a Stable State	Net Charge of Ions
IA	1	lose 1 electron ~~or gain 7 electrons~~	+1
IIA	2	lose 2 electrons ~~or gain 6 electrons~~	+2
IIIA	3	lose 3 electrons ~~or gain 5 electrons~~	+3
IVA	4	lose 4 electrons ~~or gain 4 electrons~~	+4
VA	5	~~lose 5 electrons or~~ gain 3 electrons	–3
VIA	6	~~lose 6 electrons or~~ gain 2 electrons	–2
VIIA	7	~~lose 7 electrons or~~ gain 1 electron	–1
VIIIA	8	already has the maximum number of electrons allowed in the outermost electron shell.	0

Monoatomic Ions: - ions that came from a single atom (include metal cations and non-metal anions).
- monoatomic anion ends with suffix ~ide.

Examples: Na^+ = sodium ion, Cl^- = chloride, Pb^{4+} = lead (IV) ion, Zn^{2+} = zinc ion

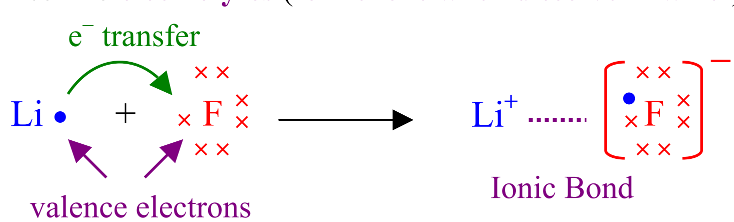
Polyatomic (Complex) Ions: - ions that contain many atoms.
- mostly anions (except NH_4^+ = ammonium ion).
- most ends with suffixes ~*ate* or ~*ite* (some ends with suffix ~*ide*).

Examples: CO_3^{2-} = carbonate, $Cr_2O_7^{2-}$ = dichromate, OH^- = hydroxide, SO_3^{2-} = sulfite

6.4: Ionic Compound

Ionic Compound: - when a metal combines with a non-metal.
- forms **ionic bonds** (electrons are "stolen" or "transfer" from one atom to another).
- dissociates into **electrolytes** (forms ions when dissolve in water)

Example: LiF



Nomenclature: - a naming system

IUPAC: - International Union of Pure and Applied Chemistry.
- an organisation that oversees the standard nomenclature of all chemicals.

Nomenclature of Ionic Compounds

1. **Balance the Cation and Anion Charges.**
2. Use **brackets** for **multiple Complex Ions (Polyatomic Ions)**.
3. When naming, use *~ide* for the **non-metal anions**.
4. **Metals** that can have **two or more different charges** must use **Roman Numerals** in the names.

Example 1: Write the chemical formulas for the following.

a. sodium chloride

Na^+ and Cl^- ⇒ **NaCl**

Need 1 Na^+ & 1 Cl^- to balance charges

b. calcium fluoride

Ca^{2+} and F^- ⇒ **CaF_2**

Need 1 Ca^{2+} & 2 F^- to balance charges

c. ammonium sulfate

NH_4^+ and SO_4^{2-} ⇒ **$(NH_4)_2SO_4$**

2 (NH_4^+) & 1 SO_4^{2-} to balance charges

d. magnesium hydroxide

Mg^{2+} and OH^- ⇒ **$Mg(OH)_2$**

Need 1 Mg^{2+} & 2 (OH^-) to balance charges

e. tin (IV) sulfite

Sn^{4+} and SO_3^{2-} ⇒ **$Sn(SO_3)_2$**

Need 1 Sn^{4+} (IV means 4+ charge) & 2 (SO_3^{2-}) to balance charges

f. aluminium oxide

Al^{3+} and O^{2-} ⇒ **Al_2O_3**

Need 2 Al^{3+} & 3 O^{2-} to balance charges

Oxyanions: - a series of polyatomic ions that contains different number of oxygen atoms.

SO_4^{2-}	sulf***ate***	ClO_4^-	***per***chlor***ate***
SO_3^{2-}	sulf***ite***	ClO_3^-	chlor***ate***
NO_3^-	nitr***ate***	ClO_2^-	chlor***ite***
NO_2^-	nitr***ite***	ClO^-	***hypo***chlor***ite***

Example 2: Name the following oxyanions.

a. BrO_4^-, BrO_3^-, BrO_2^-, BrO^-

BrO_4^-	***per***brom***ate***
BrO_3^-	brom***ate***
BrO_2^-	brom***ite***
BrO^-	***hypo***brom***ite***

b. IO_4^-, IO_3^-, IO_2^-, IO^-

IO_4^-	***per***iod***ate***
IO_3^-	iod***ate***
IO_2^-	iod***ite***
IO^-	***hypo***iod***ite***

Example 3: Name the following ionic compounds.

a. Na_2S

Na^+ = sodium
S^{2-} = sulfide
} **sodium sulfide**

b. $KMnO_4$

K^+ = potassium
MnO_4^- = permanganate
} **potassium permanganate**

c. $Li_2S_2O_3$

Li^+ = lithium
$S_2O_3^{2-}$ = thiosulfate
lithium thiosulfate

d. $CuCr_2O_7$

Cu^{2+} = copper (II)
$Cr_2O_7^{2-}$ = dichromate
} **copper (II) dichromate**

e. $Fe_3(BO_3)_2$

Fe^{2+} = iron (II)
BO_3^{3-} = borate
} **iron (II) borate**

f. $Ti(BrO_3)_3$

Ti^{3+} = titanium (III)
BrO_3^- = bromate
} **titanium (III) bromate**

Hydrate: - ionic compounds sometimes come with water molecule locked in their crystal form.
- naming contains the ionic compound name with a prefix follow by the word "*hydrate*".

Prefixes for Hydrates

1 - mono	4 - tetra	7 - hepta	10 - deca
2 - di	5 - penta	8 - octa	
3 - tri	6 - hexa	9 - nona	

Example: $CuSO_4 \bullet 5H_2O$ copper (II) sulfate pentahydrate

Assignment
6.3 pg. 145 #16, 17; pg. 146 #18, 19; pg. 148 #20, 22 and 23
6.4 pg. 151 #24, 25; pg. 153 #26, 27; pg. 155 #28, 29; pg. 156 #30 to 36

6.5: Molecular Compounds and Acids

Nomenclature of Molecular Compounds

1. Do NOT use charges to balance subscripts. Use prefixes to name or write the formula's subscripts.
2. If the first element has one atom in the molecule, do NOT use *mono~* as a prefix.
3. The last element uses the suffix *~ide*.

Prefixes for Binary Molecular Compounds

1 - mono	4 - tetra	7 - hepta	10 - deca
2 - di	5 - penta	8 - octa	
3 - tri	6 - hexa	9 - nona	

Example 1: Name the following molecular compounds.

a. CO
1 Carbon and 1 Oxygen
Carbon monoxide

b. CO_2
1 Carbon and 2 Oxygen
Carbon dioxide

c. N_2O_4
2 Nitrogen and 4 Oxygen
dinitrogen tetraoxide

Example 2: Provide the chemical formula for the following compounds

a. sulfur trioxide
1 S and 3 O ⇒ SO_3

b. diphosphorus pentaoxide
2 P and 5 O ⇒ P_2O_5

c. silicon dioxide
1 Si and 2 O ⇒ SiO_2

Common Names for Some Molecular Compounds (Memorize!)

H_2O	Water	H_2O_2	Hydrogen Peroxide	O_3	Ozone	CH_4	Methane
C_3H_8	Propane	NH_3	Ammonia	CH_3OH	Methanol	C_2H_5OH	Ethanol
$C_6H_{12}O_6$	Glucose	$C_{12}H_{22}O_{11}$	Sucrose				

Note: Do NOT use prefixes for the above common molecular compounds!

Acid: - ionic substance when dissolves in water will produce an H^+ ion.
- always in aqueous state (*aq*).

$$\boxed{H^+ + \text{Anion} \rightarrow \text{Acid}}$$

Example: $HCl_{(g)} \xrightarrow{H_2O} HCl_{(aq)}$

Nomenclature of Acid

	Ionic Compound Name		Acid Name
1.	hydrogen ~ide	→	hydro~ic acid
2.	hydrogen ~ate	→	~ic acid
3.	hydrogen ~ite	→	~ous acid

Example 3: Name the following acids.

a. $HBr_{(aq)}$

hydrogen bromide

hydrobromic acid

b. $H_2SO_{4\,(aq)}$

hydrogen sulfate

sulfuric acid

c. $H_2SO_{3\,(aq)}$

hydrogen sulfite

sulfurous acid

Example 4: Provide chemical formulas for the following acids.

a. hydrosulfuric acid

hydrosulfuric acid

hydrogen sulfide ⇒ H^+ and S^{2-}

$H_2S_{(aq)}$

b. acetic acid

acetic acid ⇒ **hydrogen acetate**

H^+ and CH_3COO^-

$HCH_3COO_{(aq)}$

c. hypochlorous acid

hypochlorous acid

hydrogen hypochlorite

⇒ H^+ and ClO^-

$HClO_{(aq)}$

6.6: Summary of Naming and Formula Writing

1. Identify whether the compound is ionic, molecular or acid. (Hint: if there is a metal or a complex ion, it is ionic).
2. Molecular compound uses prefixes. They have no charges.
3. Ionic compounds require balancing of charges. Some transition metal cations need to be specified with roman numerals.
4. Acids are originally named with "hydrogen (anion name)", and they are all in aqueous state. All acids use special rules to name depending on the suffix of the anion.

Assignment

6.5 pg. 159 #37, 38; pg. 160 #39 to 42
6.6 pg. 163 #43 and 44
Ch 6 Review: pg. 166-167 #45 to 71, 73

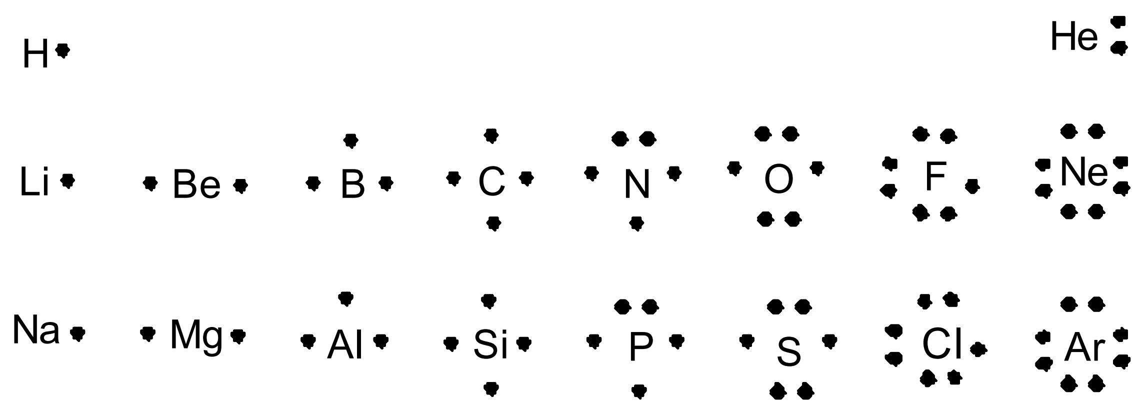
Chapter 15: Ionic Bonding and Ionic Compounds

15.1: Electron Configuration in Ionic Bonding

Valance Electrons: - electrons that are in the outermost *s* and *p* orbitals of an atom or an ion.

Lewis Structure: - sometimes refer to as **Lewis Dot Diagram** or **Electron Dot Diagram**.
- only show electrons of valence electron for all atoms involve in a molecule.

Lewis Structure for Atoms in the First Three Rows of the Periodic Table of Elements



Duet Rule: - two electrons will form a stable shell around atoms in the first period (H and He).

Octet Rule: - eight electrons are required to form a stable shell around atoms in the second and third periods.
- in general, small noble gases like Ne and Ar do not form compounds because they have already attain eight valence electrons.

Electron Configurations of Ions: - for **metals, which like to lose electrons to form cations**, they have the same electron configurations as the **noble gas of the previous row**.
- for **non-metals, which like to gain electrons to form anions**, they have the same electron configurations as the **noble gas at the end of the same row**.

Alkali Cations: - cations that were the result as alkali metals (Group 1 or IA) losing one valence electron.
- they are Li^+, Na^+, K^+, Rb^+, Cs^+ and Fr^+

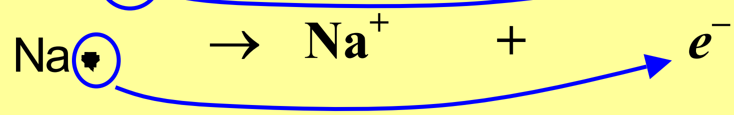
Alkaline Cations: - cations that were the result as alkaline earths (Group 2 or IIA) losing two valence electrons.
- they are Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+} and Ra^{2+}

Halide Ions: - anions that were the result as halogen (Group 17 or VIIA) gaining one valence electron to satisfy the octet rule.
- they are F^-, Cl^-, Br^-, I^- and At^-

Example 1: Write the electron configuration and an ionic equation showing the gaining or losing of electrons when each of the following atoms becomes their most popular ions. Draw the Lewis dot diagram for each atom and ion involved.

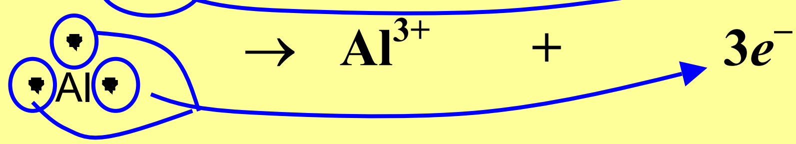
a. sodium atom to sodium ion

$$\underset{1s^2\,2s^2\,2p^6\,3s^1}{Na} \rightarrow \underset{1s^2\,2s^2\,2p^6}{Na^+} + e^-$$



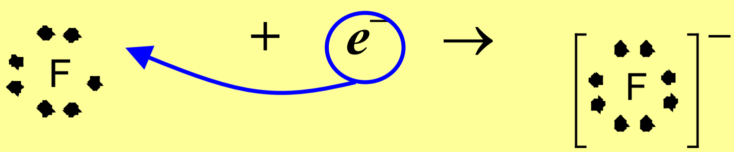
b. aluminium atom to aluminium ion

$$\underset{1s^2\,2s^2\,2p^6\,3s^2\,3p^1}{Al} \rightarrow \underset{1s^2\,2s^2\,2p^6}{Al^{3+}} + 3e^-$$



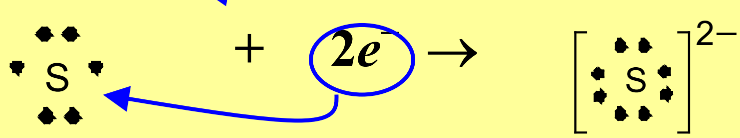
c. fluorine atom to fluoride

$$\underset{1s^2\,2s^2\,2p^5}{F} + e^- \rightarrow \underset{1s^2\,2s^2\,2p^6}{F^-}$$



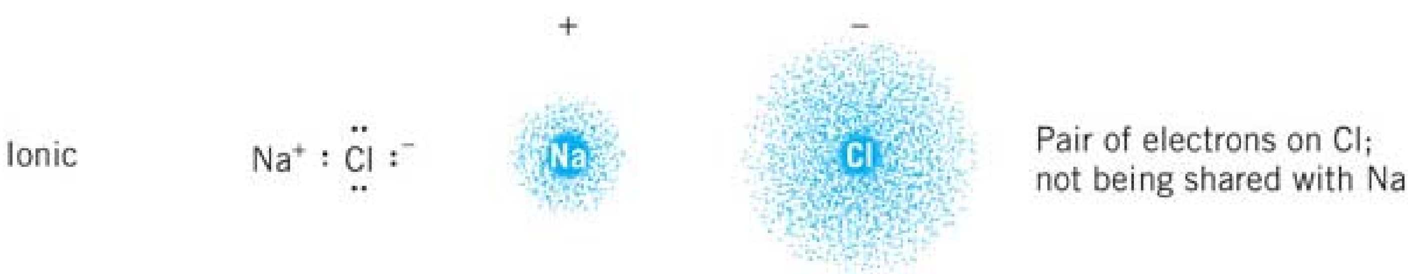
d. sulfur atom to sulfide

$$\underset{1s^2\,2s^2\,2p^6\,3s^2\,3p^4}{S} + 2e^- \rightarrow \underset{1s^2\,2s^2\,2p^6\,3s^2\,3p^6}{S^{2-}}$$



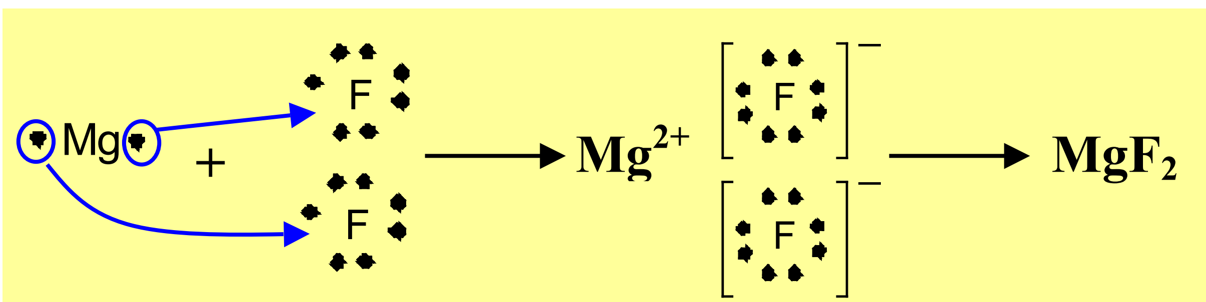
15.2: Ionic Bonds

Ionic Bonds: - the attraction force between metal and non-metal due to the transfer of electron(s) from a metal element to non-metal element.
- the resulting attraction between the cation(s) and anion(s) is the ionic bond within an ionic compound.

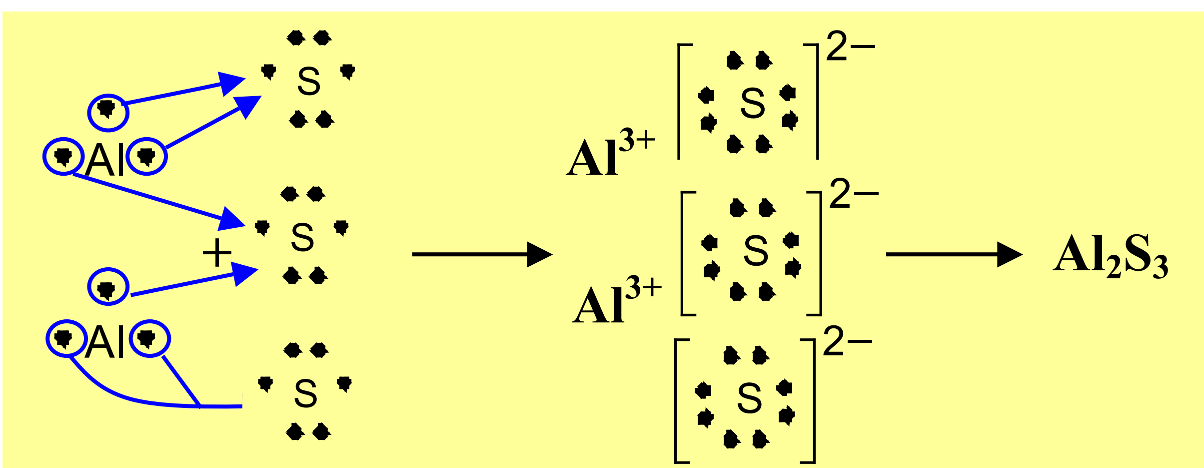


Example 2: For each ionic compound, draw the Lewis dot diagram for each atom and ion involved.

a. Mg and F

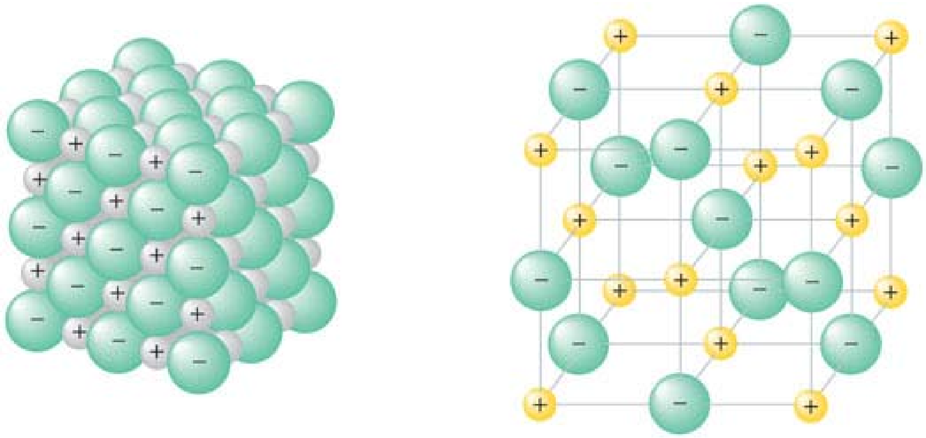


b. Al and S



Properties of Ionic Compounds

1. **Ionic Compounds have a definite Crystalline Structure and are Poor Conductors of Electricity and Heat in their Solid Form.** Conduction of electricity and heat requires ions to move freely within the solid. The lattice (crystalline) structure of the solid ionic compounds do not allow ions to move freely.



Crystalline Structure of NaCl (+ is Na^+ and – is Cl^-)

2. **Ionic solids are generally High Melting Points (typically 300°C to 1000°C).** Since a strong force can only shatter the crystal but not bend it as in metals, **the energy needed to completely break up the crystalline structure (lattice energy)** is very large and it is the same energy needed to melt the ionic compounds.

3. **Ionic solids are Hard and Brittle.** The crystalline structure of all ionic compounds holds the ions in definite positions. When the compound encountered a strong force, the close proximity of the ions stay close together. This causes the crystal to **shatter**, not bent like metal solid would.

4. **Ionic solids can be Melted to form Liquids that are Electrical Conductors.** Ionic solids melt when the ions gain enough energy to break the crystalline structure. They are move freely and can carry electrical charge through the liquid. This explains why a molten ionic substance conducts electricity, but a solid ionic material doesn't. The ions move through the liquid can carry charge from one place to another.

5. **Soluble ionic solids dissolve to form solutions that are Electrical Conductors. (Not all ionic substances are soluble in water.)** Soluble ionic compounds form **electrolytes (ions in aqueous from)** that allow the conduction of electricity.

Assignment

15.1 pg. 418 #1 to 6
15.2 pg. 421 #7 and 8; pg. 425 #9 to 13
Chapter 15 Review: pg. 432 #20 to 35, 44 to 49, 52, 54 and 55

Chapter 8: Chemical Reactions

8.1: Describing Chemical Change

Reactants: - chemicals that goes into a reaction.

Products: - chemicals that are produced from a reaction.

$$\text{Reactants} \xrightarrow{\text{“yields”}} \text{Products}$$

Chemical Word Equation: - a chemical reaction written out in words.

Skeletal Equation: - a chemical equation that does not show the relative amount of reactants and products.

States of Chemicals: - (*s*) solid, (*l*) liquid, (*g*) gas, (*aq*) aqueous – dissolved in water

Other Chemical Symbols:

1. **Heat is Added**: - $\xrightarrow{\textit{heat}}$ or $\xrightarrow{\Delta}$
2. **Catalyst is Added**: - $\xrightarrow{\textit{Name of Catalyst}}$

(**Catalyst**: - a chemical that is used to speed up a reaction but does not get consumed)

Example 1: Convert the following unbalanced chemical equations into word equations.

a. $N_{2\,(g)} + H_{2\,(g)} \rightarrow NH_{3\,(g)}$

Nitrogen gas reacts with hydrogen gas to produce ammonia gas.

nitrogen gas + hydrogen gas → ammonia gas

b. $Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$

Hydrogen gas and sodium hydroxide solution are produced when solid sodium is added to water.

sodium metal + water → sodium hydroxide solution + hydrogen gas

Example 2: Convert the following word equations into skeletal equations.

a. Heating solid diphosphorous pentaoxide decomposes into phosphorous and oxygen.

$P_2O_{5\,(s)} \rightarrow P_{4\,(s)} + O_{2\,(g)}$

Recall that phosphorus is a polyatomic element, and oxygen is a diatomic element.

b. Hypochlorous acid is neutralized by barium hydroxide solution to form water and soluble barium hypochlorite.

$HClO_{(aq)} + Ba(OH)_{2\,(aq)} \rightarrow H_2O_{(l)} + Ba(ClO)_{2\,(aq)}$

It is very important that the subscripts for these ionic compounds are correct.

Balancing Chemical Equation: - the process by which we place **coefficient (numbers in front of reactants and products)** in an attempt to equate the number of atoms or polyatomic ions for the elements or compounds in a chemical equation.

Steps involve to Balance Chemical Equation

1. Write the chemical formulas of all reactants and products with their proper subscripts and state to form a skeletal equation.
2. Count the number of atoms / polyatomic ions (always view complex ion as a group) of each chemical formula. Balance the atoms / polyatomic ions by writing the coefficient in front of the reactant / product. Do NOT mess with any of the subscripts.
3. Always balance the atoms that appear in more than two chemicals last.
4. Verify that all atoms / polyatomic ions are balanced.

Example 3: Balance the following chemical equations

a. $N_{2\,(g)} + H_{2\,(g)} \rightarrow NH_{3\,(g)}$

$N_{2\,(g)} + \underline{3}\, H_{2\,(g)} \rightarrow \underline{2}\, NH_{3\,(g)}$

b. $Al_2O_{3\,(s)} \rightarrow Al_{(s)} + O_{2\,(g)}$

$\underline{2}\, Al_2O_{3\,(s)} \rightarrow \underline{4}\, Al_{(s)} + \underline{3}\, O_{2\,(g)}$

c. $Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$

$Na_{(s)} + HOH_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$

$\underline{2}\, Na_{(s)} + \underline{2}\, HOH_{(l)} \rightarrow \underline{2}\, NaOH_{(aq)} + H_{2\,(g)}$

Sometimes, it is easier to rewrite H_2O as HOH. This is especially true when a OH^- is part of the products. The first H atom in the HOH becomes $H_{2\,(g)}$ and the remaining part, OH, becomes the polyatomic ion, OH^-.

d. $Fe(NO_3)_{3\,(aq)} + Ba(OH)_{2\,(aq)} \rightarrow Fe(OH)_{3\,(s)} + Ba(NO_3)_{2\,(aq)}$

$$\underline{2}\, Fe(NO_3)_{3\,(aq)} + \underline{3}\, Ba(OH)_{2\,(aq)} \rightarrow \underline{2}\, Fe(OH)_{3\,(s)} + \underline{3}\, Ba(NO_3)_{2\,(aq)}$$

Look at each polyatomic ion as one item. There were $(NO_3)_3$ on the left hand side and $(NO_3)_2$ on the right hand side. Hence, we need to use 2 and 3 as coefficients to balance them. Similarly, there were $(OH)_2$ on the reactant side and $(OH)_3$ on the product side. Therefore, we are required to use 3 and 2 to balance them. Note that once the coefficients are in place, the Fe and Ba atoms are also balanced.

e. $C_2H_{6\,(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{(l)}$

For burning (adding O_2) with hydrocarbons (compounds containing carbon and hydrogen), we must *balance C, H, O in that order*. This is because oxygen atoms exist in more than two compounds on the product side. After we balance carbon and hydrogen, we have the following number of oxygen atoms.

$C_2H_{6\,(g)} + \_\_\, O_{2\,(g)} \rightarrow \underline{2}\, CO_{2\,(g)} + 3\, H_2O_{(l)}$

need 7 oxygen 4 oxygen 3 oxygen

$C_2H_{6\,(g)} + \underline{\tfrac{7}{2}}\, O_{2\,(g)} \rightarrow \underline{2}\, CO_{2\,(g)} + \underline{3}\, H_2O_{(l)}$

Multiply all coefficients by 2:

$2\, C_2H_{6\,(g)} + \underline{7}\, O_{2\,(g)} \rightarrow \underline{4}\, CO_{2\,(g)} + \underline{6}\, H_2O_{(l)}$

Since elemental oxygen is diatomic, the 7 oxygen needed on the reactant side needs to have a coefficient of 7/2.

This can easily be converted to a whole number by multiplying all the coefficients by 2.

Example 4: Rewrite the following word equations into skeletal equations and balance them to form chemical equations.

a. Sulfur is burned in the presence of oxygen to form sulfur dioxide gas.

First, we write the skeletal equation.
(Recall that sulfur is a polyatomic element, and oxygen is a diatomic element.)

$$S_{8\,(s)} + O_{2\,(g)} \rightarrow SO_{2\,(g)}$$

Then, we balance the equation.

$$S_{8\,(s)} + \underline{8}\ O_{2\,(g)} \rightarrow \underline{8}\ SO_{2\,(g)}$$

b. Hydrogen gas is added to propyne (C_3H_4) gas with the aid of a platinum catalyst to form propane gas.

First, we write the skeletal equation.
(Propane is C_3H_8 – a formula that you should have memorized.)

$$C_3H_{4\,(g)} + H_{2\,(g)} \xrightarrow{Pt} C_3H_{8\,(g)}$$

Then, we balance the equation.

$$\underbrace{C_3H_{4\,(g)}}_{\text{4 hydrogen}} + \underbrace{\underline{2}\ H_{2\,(g)}}_{\textit{need 4 more hydrogen}} \xrightarrow{Pt} \underbrace{C_3H_{8\,(g)}}_{\text{8 hydrogen}}$$

8.2: Types of Chemical Reactions

There are 5 basic types of chemical reactions:

1. **Formation or Composition** **(Element + Element → Compound)**

   Example: $2\ Mg_{(s)} + O_{2\,(g)} \rightarrow 2\ MgO_{(s)}$

2. **Deformation or Decomposition** **(Compound → Element + Element)**

   Example: $2\ Al_2O_{3\,(s)} \rightarrow 4\ Al_{(s)} + 3\ O_{2\,(g)}$

3. **Single Replacement** **(Element + Compound → Element + Compound)**

   Example: $2\ AgNO_{3\,(aq)} + Cu_{(s)} \rightarrow 2\ Ag_{(s)} + Cu(NO_3)_{2\,(aq)}$

4. **Double Replacement** **(Compound + Compound → Compound + Compound)**

   Example: $AgNO_{3\,(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3\,(aq)}$

5. **Hydrocarbon Combustion** **(Hydrocarbon + Oxygen → Carbon Dioxide + Water)**

   Example: $CH_{4\,(g)} + 2\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(g)}$

Precipitation Reaction: - a reaction where a precipitate (new solid) is formed as a product.

Neutralization Reaction: - a reaction between an acid and a base where water is formed as a product.

To Predict Products and Balance Chemical Equations:

1. Write the correct chemical formulas for all products and reactants with proper subscripts. The presence of metals or ionic compounds indicates that we will need to use ions and charges to form any products.
2. For hydrocarbon combustion, balance in the order of C, H, and then O.
3. For other type of reactions, balance the equation for each type of cations and anions. Do NOT break up complex ions. Water may be written as HOH (H^+ and OH^-) in single and double replacement reactions.
4. Check with the Solubility Table (see Section 18.1) and the Table of Elements for the states of chemicals.

Example 1: Predict the product(s) along with the states, indicate the type of reaction, and balance the following chemical reactions.

a. Sulfur trioxide gas is produced from its elements.

Formation: $S_{8\,(s)} + \underline{12}\ O_{2\,(g)} \rightarrow \underline{8}\ SO_{3\,(g)}$

b. A solid piece of zinc is immersed in an iron (III) chloride solution.

Single Replacement: $\underline{3}\ Zn_{(s)} + \underline{2}\ FeCl_{3\,(aq)} \rightarrow \underline{3}\ ZnCl_{2\,(aq)} + \underline{2}\ Fe_{(s)}$

Zn^{2+} Fe^{3+} Cl^-

c.

Hydrocarbon Combustion: $C_3H_{8\,(g)} + \underline{5}\ O_{2\,(g)} \rightarrow \underline{3}\ CO_{2\,(g)} + \underline{4}\ H_2O_{(g)}$

d. Chlorine gas is bubbled through a copper (II) iodide solution.

Single Replacement: $Cl_{2\,(g)} + CuI_{2\,(aq)} \rightarrow CuCl_{2\,(aq)} + I_{2\,(s)}$

Cl^- Cu^{2+} I^-

e. Ammonia gas is decomposed into its elements.

Decomposition: $\underline{2}\ NH_{3\,(g)} \rightarrow N_{2\,(g)} + \underline{3}\ H_{2\,(g)}$

f. Sulfuric acid is neutralized by sodium hydroxide solution.

Double Replacement: $H_2SO_{4\,(aq)} + \underline{2}\ NaOH_{(aq)} \rightarrow 2\ HOH_{(l)} + Na_2SO_{4\,(aq)}$

H^+ SO_4^{2-} Na^+ OH^-

g. Propanol ($C_3H_7OH_{(l)}$) is accidentally ignited.

Hydrocarbon Combustion: $C_3H_7OH_{(l)} + \frac{9}{2}\ O_{2\,(g)} \rightarrow \underline{3}\ CO_{2\,(g)} + \underline{4}\ H_2O_{(g)}$

(Multiply Coefficients by 2) $\underline{2}\ C_3H_7OH_{(l)} + \underline{9}\ O_{2\,(g)} \rightarrow \underline{6}\ CO_{2\,(g)} + \underline{8}\ H_2O_{(g)}$

h. Lead (II) nitrate solution is reacted with chromium (III) sulfate solution.

Double Replacement: $\underline{3}\ Pb(NO_3)_{2\,(aq)} + Cr_2(SO_4)_{3\,(aq)} \rightarrow \underline{2}\ Cr(NO_3)_{3\,(aq)} + \underline{3}\ PbSO_{4\,(s)}$

Pb^{2+} NO_3^- Cr^{3+} SO_4^{2-}

i. Octane ($C_8H_{18\,(l)}$) is combusted in an automobile.

Hydrocarbon Combustion: $C_8H_{18\,(l)} + \frac{25}{2} O_{2\,(g)} \rightarrow \underline{8}\ CO_{2\,(g)} + \underline{9}\ H_2O_{(g)}$

(Multiply Coefficients by 2) $\underline{2}\ C_8H_{18\,(l)} + \underline{25}\ O_{2\,(g)} \rightarrow \underline{16}\ CO_{2\,(g)} + \underline{18}\ H_2O_{(g)}$

Assignment

8.1 pg. 206 #1, 2; pg. 209 #3, 4; pg. 210 #5 to 7; pg. 211 #8 to 11

8.2 pg. 214 #13, 14; pg. 216 #15, 16; pg. 218 #17;
pg. 220–221 #18 to 21; pg.224 #22, 23 (omit 23c)

8.3: Reactions in Aqueous Solution

Molecular Equation: - a chemical equation where compounds are written in their chemical formulas.

Complete Ionic Equation: - a chemical equation where all compounds that are soluble are written in the ionic components (**slightly soluble compounds are not separated into ions**).

Net Ionic Equation: - an ionic equation that only shows the ions responsible in forming the precipitate. **Spectator Ions** (ions that do not form the precipitate) are omitted.

Example 1: Predict all products form when an ammonium phosphate solution reacts with a calcium chloride solution. Explain the reaction in a form of a balanced

a. Molecular Equation

$$2\ (NH_4)_3PO_{4\,(aq)} + 3\ CaCl_{2\,(aq)} \rightarrow 6\ NH_4Cl_{(aq)} + Ca_3(PO_4)_{2\,(s)} \downarrow$$

NH_4^+ PO_4^{3-} Ca^{2+} Cl^- **(precipitate)**

b. Complete Ionic Equation

$$6\ NH_4^+{}_{(aq)} + 2\ PO_4^{3-}{}_{(aq)} + 3\ Ca^{2+}{}_{(aq)} + 6\ Cl^-{}_{(aq)} \rightarrow 6\ NH_4^+{}_{(aq)} + 6\ Cl^-{}_{(aq)} + Ca_3(PO_4)_{2\,(s)} \downarrow$$

(Precipitate does NOT separate into ions)

c. Net Ionic Equation

$$2\ PO_4^{3-}{}_{(aq)} + 3\ Ca^{2+}{}_{(aq)} \rightarrow Ca_3(PO_4)_{2\,(s)}$$

(Only write the ions that contribute to the precipitated chemical species)

Example 2: Predict all products form when sulfuric acid reacts with a lithium hydroxide solution. Explain the reaction in a form of a balanced

a. Molecular Equation

$$H_2SO_{4\,(aq)} + 2\ LiOH_{(aq)} \rightarrow Li_2SO_{4\,(aq)} + 2\ HOH_{(l)}$$

H^+ SO_4^{2-} Li^+ OH^- **(liquid water)**

b. Complete Ionic Equation

$$2\ H^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\ Li^+{}_{(aq)} + 2\ OH^-{}_{(aq)} \rightarrow 2\ Li^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2\ HOH_{(l)}$$

(Pure Liquid does NOT separate into ions)

c. Net Ionic Equation

$$2\ H^{+}_{(aq)} + 2\ OH^{-}_{(aq)} \rightarrow 2\ HOH_{(l)}$$
$$H^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow H_2O_{(l)}$$

This is the main result of acid-base neutralization (the formation of water).

(Only write the ions that contribute to the pure liquid species)

Example 3: Predict all products form when solid aluminum reacts with a copper (II) nitrate solution. Explain the reaction in a form of a balanced

a. Molecular Equation

$$2\ Al_{(s)} + 3\ Cu(NO_3)_{2\ (aq)} \rightarrow 2\ Al(NO_3)_{3\ (aq)} + 3\ Cu_{(s)} \downarrow$$

Al^{3+} Cu^{2+} NO_3^{-} **(precipitate)**

b. Complete Ionic Equation

$$2\ Al_{(s)} + 3\ Cu^{2+}_{(aq)} + 6\ NO_3^{-}{}_{(aq)} \rightarrow 2\ Al^{3+}_{(aq)} + 6\ NO_3^{-}{}_{(aq)} + 3\ Cu_{(s)} \downarrow$$

c. Net Ionic Equation

$$2\ Al_{(s)} + 3\ Cu^{2+}_{(aq)} \rightarrow 2\ Al^{3+}_{(aq)} + 3\ Cu_{(s)}$$

(Need to write all the ions on both sides that correspond to any solid used or formed)

Assignment

8.3 pg. 226 #25, pg. 228 #26 to 31

Chapter 8 Review: pg. 232 #32 to 36, 38 to 53 (omit 43e, 46 & 52c)