Name ____________________ **AP CHEM** **___/___/___**

Chapter 2 Outline
Atoms, Molecules & Ions

Fundamental Chemical Laws

- Antoine Lavoisier (1743-1794)
  - Father of Modern Chemistry
  - Explained the nature of combustion; showed combustion involves oxygen, not phlogiston
  - His experiments suggested that mass is not created or destroyed (Law of Conservation of Mass) **(LO 3.6)**

- Joseph Proust (1754-1826)
  - Showed that a given compound contains exactly the same proportion of elements by mass.
    - Example: Copper(II) Carbonate has a definite mass ratio of Cu: 5.3 parts to 4 parts oxygen to 1 part carbon; $CuCO_3$, g.f.m = (Cu, 63.5 + C, 12.0 + O, 48.0 = 123.5), composition = (Cu, 63.5/123.5 = 0.514; C, 12.0/123.5 = 0.0972; O, 48.0/123.5 = 0.389), ratio = (Cu, 0.514/0.0972 = 5.29, C, 0.0972/0.0972 = 1.00, O = 0.389/0.0972 = 4.00)
  - Proust's Law, a.k.a. Law of Definite Proportion **(LO 1.1)**
- John Dalton (1766-1844) **(LO 3.6)**
  - Proust's work inspired Dalton to think about atoms as parts of the elements
  - His reasoning was that if elements were composed of tiny individual parts(atoms) then a given compound should always contain the same combination of these atoms
  - Discovered that when 2 elements form a series of compounds the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers – Law of Multiple Proportions
    - **Example 1:**

	Mass of Oxygen That Combines with 1 gram of carbon	Chemical Formula
Compound 1	1.33 grams	CO
Compound 2	2.66 grams	CO_2

Since we now know the mass of carbon is 12.01 amu and the mass of oxygen is 16.00 amu we can understand the formula of the first compound to be **CO (12.01 x 1.33 = 16.0)**. In the second compound there is twice as much mass of oxygen per gram of carbon so we can determine the formula to be **CO_2**.

Dalton's Atomic Theory (LO 1.13)

- 1808 published – A New System of Chemical Philosophy
  - Each element is made up of atoms
  - Atoms of a given element are identical; atoms of different elements are different
  - Compounds are formed when atoms of different elements combine with one another. A given compound always has the same relative numbers and types of atoms.
  - Atoms are not changed in chemical reactions.
- Designed first table of atomic masses, though most of his masses were wrong because of his incorrect assumptions about formulas
- Joseph Gay-Lussac (1778-1850) **(LO 3.4, 3.6)**
  - In 1809 performed experiments in which he measured the volumes of gases that reacted with each other (at constant temperature and pressure)
    - Example: 2 volumes of hydrogen react with 1 volume of oxygen to form 2 volumes of water vapor ($2H_2 + O_2 \rightarrow 2H_2O$)

- Amadeo Avogadro (1776-1856) **(LO 3.4, 3.6)**
    - In 1811 Avogadro interpreted these results by proposing that at the same temperature and pressure, equal volumes of different gasses contain the same number of particles (Avogadro's hypothesis)
    - If Avogadro is correct, the Gay-Lussac's result can be expressed as: 2 molecules of hydrogen react with 1 molecule of oxygen to form 2 molecules of water

Early Experiments to Characterize the Atom

- The electron **(LO 1.7)**
    - J.J. Thomson
        - English physicist who studied electrical discharges in cathode-ray tubes
            - Cathode ray tubes produce a ray of light from the cathode end of the tube
        - Thomson postulated that the ray was a stream of negative charged particles, now called electrons
        - Charge-to-mass ratio = -1.76×10^{8} C/g (Coulombs per gram)
        - Since atoms are electrically neutral, Thomson assumed that atoms must contain a positively charged particle to balance the charge of the electron.
        - Plum Pudding Model (J.J. Thomson is given credit for this model but the idea was first suggested by English mathematician and physicist William Thomson(no relation) a.k.a. Lord Kelvin)
    - Robert Millikan (1868-1953)
        - Determined the magnitude of the electron charge.
        - With the magnitude of the electron charge and the charge to mass ratio, Millikan was able to calculate the mass of the electron to be 9.11×10^{-31} kg
- Radioactivity
    - 1896 Henri Becquerel found accidentally that a piece of a mineral containing uranium could produce its image on a photographic plate in the absence of light
    - Becquerel attributed this phenomenon to a spontaneous emission of radiation by uranium, called radiation
    - Three types of radioactive emission: gamma (γ) rays, beta (β) particles and alpha (α) particles
        - γ rays – high energy light
        - β particles – high speed electron
        - α particles – helium nuclei (a particle with a 2+ charge), 7300 times the mass of an electron
- Nuclear Atom **(LO 1.13)**
    - In 1911 Ernest Rutherford carried out experiments to test Thomson's plum pudding model
    - Rutherford shot alpha particles at a thin sheet of metal foil. He expected the alpha particles to travel through the gold foil
    - Most alpha particles passed straight through the gold foil but some were deflected at large angles and a few were deflected straight back
    - Rutherford said, "It's about as credible as shooting a 12" shell at a piece of tissue paper and having it come back and hit you."
    - Rutherford determined that the atom was mostly empty space and that at the core there is a small dense nucleus (Latin: little nut).

The Modern View of Atomic Structure

Subatomic Particle	Symbol	Location	Charge
proton	p^{+}	nucleus	+1
neutron	n^{0}	nucleus	0
electron	e^{-}	outside nucleus	-1

- **The chemistry of an atom results from its electrons.**
- The tiny nucleus accounts for almost all of the atom's mass (A nucleus the size of a pea would has a mass of 250 million tons (500 billion pounds.).

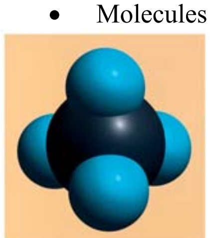
- Isotopes – atoms of an element with the same number of protons but different numbers of neutrons. **(LO 1.13, 1.14)**
- Because the chemistry of an atom is due to its electrons, isotopes show almost identical chemical properties.
- Atomic number(Z) = # of protons = # electrons (in a neutral atom only)
- Mass number(A) = # protons + # neutrons

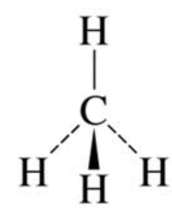
- **For example**: Two naturally occurring isotopes of chlorine are chlorine-35 & chlorine-37. Thirty-five and thirty-seven are the mass numbers for the two isotopes. Both isotopes have the same atomic number, number of protons and electrons.

isotope name	atomic number (Z)	# protons	# neutrons	mass number (A)	# electrons
chlorine-35	17	17	18	35	17
chlorine-37	17	17	20	37	17

Molecules & Ions

- The forces that hold atoms together in compounds is called chemical bonds.
- **Covalent Bonds** – atoms ***share*** electrons to form molecules.
- Molecules are represented by chemical formulas.





Methane

- **Structural formulas** give more information by indicating the individual bonds (shown with lines). The structural formula of methane, CH_4, is shown to the right. **(LO 2.21, 2.31, 2.32)**
- A **space-filling model** shows the relative sizes of the atoms as well as their relative orientation in the molecule. The space-filling model of methane is shown to the left. **(LO 2.21, 2.31, 2.32)**

- **Ball-and-stick models** are also used to represent molecules. The ball-and-stick model of methane is represented to the right.**(LO 2.21, 2.31, 2.32)**
- Ionic Bonds occur between a cation and an anion. A cation is a positively charged ion that has lost an electron. An anion is a negatively charged ion that has gained an electron. **(LO 2.23, 2.24)**
- Because anions and cations have opposite charges, they attract each other. This is referred to as ionic bonding.
- A solid consisting of oppositely charged ions is called an ionic solid, or a salt.

An Introduction to the Perioidic Table (LO 1.9)

- Most elements are metals. Metals are efficient conductors of heat and electricity, malleable, ductile and often lustrous. Metals tend to lose electrons when they form ions.
- There are few nonmetals. Nonmetals tend to gain electrons and acquire a negative charge when they form ions.
- Elements in the same vertical column are called groups or families and have similar chemical properties.
- Group 1 elements, (with the exception of hydrogen), are called the alkali metals and all have one valence electron and a 1+ charge.
- Group 2 elements are called alkaline-earth elements and all have 2 valence electrons and a 2+ charge.
- Groups 3-12 elements are called transitions metals and most have 2 valence electrons. Their charges vary.
- Group 13 elements are part of the boron family and all have 3 valence electrons. Most have a 3+ charge.
- Group 14 elements are part of the carbon family (crystallogens) and all have 4 valence electrons. Most have a 4+ charge. Carbon can also be 4-. Some metals can also be 2+.

- Group 15 elements are part of the nitrogen family (pnictogens) and all have 5 valence electrons. Non-metals have a 3- charge; metals vary in charge.
- Group 16 elements are part of the oxygen family and all have 6 valence electrons. Non-metals have a 2- charge; metals vary in charge.
- Group 17 elements are called halogens and all have 7 valence electrons and a 1- charge.
- Group 18, are called the noble gases. These elements have 8 valence electrons and are inert(non-reactive) and have no charge.
- The "charges" mentioned above refer to the most common ion formed when these elements form ionic compounds. Elements are neutral until they gain, lose or share their electrons and form compounds.
- Horizontal rows of elements on the periodic table are called periods.
- Vertical columns on the periodic table are called groups or families because they have the same number of valence electrons and thus similar properties.

Naming Simple Compounds

- An **ionic compound** is a compound that is formed between a **metal** and a **non-metal**. (Metalloids can also be used in ionic compounds, sometimes as cations and sometimes as anions, depending on the properties of the specific element.). In ionic compounds the **metal will always be a cation** and the **non-metal will always be an anion**. Please note**, the negative oxidation numbers we wrote on top of Groups 14, 15, 16 & 17 on our periodic tables refer only to the non-metals and the metalloids.** The metals in these columns have different oxidation numbers.
- When forming ionic compounds the goal is to balance the number of positive charges with the number of negative charges. More specifically, you want to ensure that the number of electrons that the cations are giving up is equal to the number of electrons the anions need so that both have full outer energy level.
- Binary ionic compounds contain only two elements, one is the cation and the other is the anion. It is important to remember that when writing binary ionic compounds **THE CATION MUST ALWAYS BE WRITTEN FIRST**. The rest of the rules will be outlined in the following example.

Part I. How to Write a Binary Ionic Compound Formula

Example: Write the formula for the compound between barium and sulfur.

Description of Action	Action
1. Write the symbol of the **cation** with its charge.	**1.** Ba^{2+}
2. To the right of the cation, write the **anion** and its charge.	**2.** Ba^{2+} S^{2-}
3. **Cross** each element's oxidation number to the **lower right side** of the other element's symbol.	**3.** Ba^{2+} S^{2-} **Result:** Ba_{2-} S_{2+}
4. Remove all (+) signs, (-) signs and ones.	**4.** Ba_2S_2
5. Reduce, if necessary. You can only reduce when the subscripts of all the symbols have a common denominator.	**5.** Since Ba and S both have a 2 for a subscript, it can be reduced to 1. And, since ones are not written, the answer is: **BaS**

Part II. Naming Binary Ionic Compounds

- IMPORTANT: **On your periodic table, cross out the suffixes for the non-metals and metalloids that form anions. For each of the halogens, cross out –ine; for selenium and tellurium, cross out –ium; for carbon, arsenic, and sulfur, cross out the last two letters of each name; and for oxygen, nitrogen and phosphorus, cross out the last four letters of each element's name.** What remains will be referred to as the anion's root name. We will use this root name and the new suffix **–ide** when naming ionic compounds.

Example: Name the compound we made earlier using barium and sulfur, **BaS**.

Description of Action	Action
1. Write the name of the cation.	**1.** barium
2. To the right of the cation name, write anion's root name.	**2.** barium sulf
3. Add the suffix **–ide** to the end of the anion.	**3.** barium sulfide

Sometimes groups of elements will combine such that they have an overall charge. These are known as polyatomic ions. The chart below lists polyatomic ions and their names. All will be used at different times in class, however, only the highlighted ones need to be MEMORIZED for tests and quizzes. I suggest making flash cards and practicing the names and formulas.

•SYMBOLS OF COMMON POLYATOMIC IONS•					
$(AsO_3)^{3-}$	arsenite	$(C_2O_4)^{2-}$	oxalate	$(N_3)^{1-}$	azide
$(AsO_4)^{3-}$	arsenate	$(CrO_4)^{2-}$	chromate	$(NH_2)^{1-}$	amide
$(BO_3)^{3-}$	borate	$(Cr_2O_7)^{2-}$	dichromate	$(NH_4)^{1+}$	AMMONIUM
$(B_4O_7)^{2-}$	tetraborate	$(HCO_3)^{1-}$	bicarbonate	$(NO_2)^{1-}$	nitrite
$(BrO)^{1-}$	hypobromite	$(HC_2O_4)^{1-}$	bioxalate	$(NO_3)^{1-}$	nitrate
$(BrO_3)^{1-}$	bromate	$(H_3O)^{1+}$	HYDRONIUM	$(O_2)^{2-}$	peroxide
$(CHO_2)^{1-}$	formate	$(HPO_4)^{2-}$	biphosphate	$(OH)^{1-}$	hydroxide
$(C_2H_3O_2)^{1-}$	acetate	$(H_2PO_4)^{1-}$	dihydrogen phosphate	$(PO_3)^{3-}$	phosphite
$(C_4H_4O_6)^{1-}$	tartrate	$(HS)^{1-}$	bisulfide	$(PO_4)^{3-}$	phosphate
$(C_6H_5O_7)^{3-}$	citrate	$(HSO_3)^{1-}$	bisulfite	$(SCN)^{1-}$	thiocyanate
$(ClO)^{1-}$	hypochlorite	$(HSO_4)^{1-}$	bisulfate	$(SO_3)^{2-}$	sulfite
$(ClO_2)^{1-}$	chlorite	$(IO)^{1-}$	hypoiodite	$(SO_4)^{2-}$	sulfate
$(ClO_3)^{1-}$	chlorate	$(IO_2)^{1-}$	iodite	$(S_2O_3)^{2-}$	thiosulfate
$(ClO_4)^{1-}$	perchlorate	$(IO_3)^{1-}$	iodate	$(SeO_4)^{2-}$	selenate
$(CN)^{1-}$	cyanide	$(IO_4)^{1-}$	periodate	$(SiF_6)^{2-}$	hexafluorosilicate
$(CO_3)^{2-}$	carbonate	$(MnO_4)^{1-}$	permanganate	$(SiO_3)^{2-}$	silicate
The word hydrogen can be substituted for the prefix bi-. (i.e. hydrogen sulfide = bisulfide)					

These polyatomic ions are treated just like any other ion when writing chemical formulas. As above, you would cross the charges.

Part III: How to Write a Formula for an Ionic Compound that has a Polyatomic Ion
Example: Write the formula for the ionic compound formed between **aluminum** and **phosphite**.

Description of Action	Action
1. Write the symbol of the cation with its charge.	**1.** Al^{3+}
2. To the right of the cation, write the polyatomic anion and its charge.	**2.** Al^{3+} $(PO_3)^{3-}$
3. Cross each element's oxidation number to the lower right side of the other element's symbol.	**3.** Al^{3+} $(PO_3)^{3-}$ **Result:** Al_{3-} $(PO_3)_{3+}$
4. Remove all (+) signs, (-) signs and ones.	**4.** $Al_3 (PO_3)_3$
5. Reduce, if necessary. Remember, **do not touch anything inside the parenthesis**.	**5.** $Al (PO_3)$
6. If there is no subscript outside the anion's parenthesis, remove the parenthesis.	**6. Answer:** $Al PO_3$

Part IV: Naming Ionic Compounds that have Polyatomic Ions
Example: Name the compound: **Al PO_3**

Description of Action	Action
1. Write the name of the cation.	**1.** aluminum
2. To the right of the cation name, write anion's name.	**2.** aluminum phosphite

Many transition metals form ions with different charges. You will need to **MEMORIZE** the charges of the following transition metals. Others transition metals also have multiple charges but these are the ones you will be quizzed/tested on. Many can be figured out with a good understanding of the periodic table.

Metal	Charges	Metal	Charges
Sc	3+	Cu	1+, 2+
Ti	3+, 4+	Zn	2+
Cr	2+, 3+, 6+	Ag	1+
Mn	2+, 3+, 4+, 6+, 7+	Au	1+, 3+
Fe	2+, 3+	Hg	1+, 2+ *note mercury(I) is $(Hg_2)^{2+}$
Co	2+, 3+	Sn	2+, 4+
Ni	2+, 3+	Pb	2+, 4+

Part V: How to Write a Formula for an Ionic Compound of a Multivalent Transition Metal
Example: Write the formula for copper(II) chloride.

Description of Action	Action
1. Write the symbol for the given cation name.	**1.** Cu
2. Write the number in parenthesis as the cation's charge.	**2.** Cu^{2+}
3. To the right of the cation, write the anion.	**3.** Cu^{2+} Cl^{1-}
4. Cross each element's oxidation number to the lower right side of the other element's symbol.	**4.** Cu^{2+} Cl^{1-} **Result:** Cu_{1-} Cl_{2+}
5. Remove all (+) signs, (-) signs and ones.	**5.** Cu Cl_2
6. Reduce if necessary. Remember, if you are using a polyatomic ion, DO NOT change anything in the parenthesis.	**6. Not necessary:** Cu Cl_2
7. If you are using a polyatomic ion and there is no number outside of the parenthesis, you can remove the parenthesis.	**7. No polyatomic ions:** $CuCl_2$

Part VI: Naming Ionic Compounds with Multivalent Cations (Note this is just one method – other ways to do this may be easier but more difficult to explain in a worksheet.)
Example: Name **$FeSO_4$**

Description of Action	Action
1. Name the cation.	**1. iron**
2. If the element can have more than one charge, write empty parenthesis after the cation's name.	**2.** Yes, iron has charges of 2+ or 3+ **iron( )**
3. Name the anion. Leave the parenthesis blank.	**3. iron( ) sulfate**
4. If they are not already written, put parenthesis around any polyatomic ions.	**4.** Fe (SO_4) Sulfate is polyatomic, so I put parenthesis around it. Iron is not polyatomic, so it does not need parenthesis.
5. Write the anion's charge to the top right of its symbol, outside of the parenthesis	**5.** Fe $(SO_4)^{2-}$
6. Multiply the anion's **charge** and the anion's **subscript**. If the anion is polyatomic, use the subscript outside of the parenthesis. If there is no number written, we can assume it is one.	**6.** For this formula we would **multiply** 2- (charge) x 1 (subscript). 2 x 1 = **2**
7. Divide the **result** by the **subscript** of the cation. Again, if there is no number written, assume the subscript is one.	**7.** Our result was (2) and there is no subscript for Fe, so we would divide: 2 ÷ 1 = **2**
8. Your new result is the roman numeral to put in parenthesis after the cation's name.	**8. iron(II) sulfate**

Part VII: Naming Covalent Compounds

- The system of naming binary compounds of **TWO (2) NON-METALS** does not really have an officially accepted name, but it is often called the Greek system (or method). It involves use of Greek prefixes when naming binary compounds formed between two nonmetals. The prefixes are listed below.

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

Example: Name the compound with the formula: N_2O_5

Description of Action	Action & Explanation
1. Identify the prefix that corresponds to the subscript following the first symbol. **NEVER USE THE PREFIX MONO- BEFORE THE FIRST ELEMENT NAME!**	**1.** di There are 2 nitrogen so we have to use the prefix "di-".
2. Add the name of the first element to the end of the prefix.	**2.** dinitrogen
3. Write the prefix for the subscript that follows the second element. You must leave a space between the first name and the second name.	**3.** dinitrogen penta There are 5 oxygen so we must use the prefix penta.
4. Attach the root name of the second element to the second prefix.	**4.** dinitrogen pentaox
5. Add "-ide" to the end of the second element's root name.	**5.** dinitrogen pentaoxide

Part VIII: Writing formulas for Covalent Compounds

Example: Write the formula for dinitrogen trioxide.

Description of Action	Action & Explanation
1. Look at the first name of the compound. Identify the element name. Write the symbol for this element.	**1.** N In dinitrogen, the elements name is nitrogen. Nitrogen's symbol is N.
2. If the first name of the compound has a prefix, write the number the prefix refers to as the symbol's subscript.	**2.** N_2 We have **di**nitrogen. "Di-" means two, so I wrote a two after N.
3. Look at the second name of the compound and identify the element root name. Write the symbol for the root name.	**3.** N_2O The second name of this compound is trioxide. There is an "ox-" in there! "Ox" refers to oxygen. Oxygen's symbol is O. So, I write that O that you see above.
4. Determine the number that the prefix of the second name refers to and write this number as the second symbol's subscript.	**4.** N_2O_3 The second name is trioxide. "Tri-" means 3. So, I wrote a 3 after the O.

Part IX: How to Name an Acid

- An **acid** is a compound that produces hydrogen ions when dissolved in water. Acids are named according to how their ionic compound name ends.
- If the anion does NOT contain oxygen, the compound is named: ***hydro-(anion root)-ic acid***
  Example: Name: HCl
  Ionic name: hydrogen chloride, Acid name: **hydrochloric acid**
- If the ionic compound name ends in "**-ate**", the compound is named: **(*anion root)-ic acid***
  Example: Name: $HClO_3$
  Ionic name: hydrogen chlorate, Acid name: **chloric acid**
- If the ionic compound name ends in "**-ite**", the compound is named: ***(anion root)-ous acid***
  Example: H_2SO_3
  Ionic name: hydrogen sulfite, Acid name: **sulfurous acid**

Part X: Writing Formulas for Acids

When writing the formula for an acid, work backwards. Non-organic acids(covered in chapter 14) always begin with Hydrogen. Cross the charges using H^+ and either the monoatomic anion or the polyatomic anion in the name.

Example: Write the formula for nitric acid.

Description of Action	Action & Explanation
1. Look at the first name of the compound. Identify the ionic compound name.	**1.** "Nitric acid" doesn't contain "hydro" so it means "nitric" comes from the polyatomic ion "nitrate". (-ic is from –ate and –ous is from –ite).
2. All non-organic acids begin with the cation H^+. Write H^+ as your cation.	**2.** H^+
3. Add the anion.	**3.** H^+ $(NO_3)^{1-}$
4. Cross charges and reduce if necessary.	**4.** HNO_3