

AP Chemistry Summer Packet

AP Chemistry student-

You are receiving this packet because you are enrolled in Advanced Placement Chemistry for the upcoming school year. I applaud you for taking on the challenge of AP Chemistry. This course is much more challenging than General Chemistry but I will be there every step of the way to support you. The students whom have completed this course have felt better prepared for college science courses and have a deeper understanding of the inner workings of the world around them. We will perform many college level labs and as such you will have higher expectations and be able to perform more complex activities. If you have any questions about the expectations or the class, you can email at the address below. The goal of this packet is to review concepts that were taught during your first year of Chemistry and introduce you to the AP Chemistry course. It is expected that on day one these are skills that you have mastered and remember well. This packet will help you keep the information fresh over the summer. This is not a required packet but those that complete it are more prepared for the class. I have provided a great deal of reminders and information which are each accompanied by a few problems to allow you to practice. I have compiled the practice problems in a separate document if you wish to print these. If you have any questions, you may use the internet for support or email me at mkmccarty@bluevalleyk12.org. I, also, encourage you to join our AP Chemistry group at www.Edmodo.com as soon as possible. **Our group code is: fcms3.** I will post a partial key to the packet later in the summer. This is, also, a great forum to ask questions because everyone can benefit from the answers. I encourage you to answer each other's questions as well. While you will not be graded on this packet you will be given quizzes and/or an exam over the material within the first month of school. Congratulations on taking on the challenge of AP chemistry and I look forward to seeing you in class! Good luck on the packet and have a wonderful summer. ☺

Thank you,

Mrs. McCarty

Review and know ALL Polyatomics. You will have a quiz Week 1 when you return for school

** Highly important for the rest of the year!! You will see these a lot**

Per__ate	__ate	__ate name	__ite	hypo__ite	__ide (for comparison)
	PO_4^{3-}	phosphate	PO_3^{3-} phosphite		P^{3-} phosphide
	SO_4^{2-}	sulfate	SO_3^{2-} sulfite	SO_2^{2-} hyposulfite	S^{2-} sulfide
	CrO_4^{2-}	chromate			
	CO_3^{2-}	carbonate			
	$\text{C}_2\text{O}_4^{2-}$	oxalate			
	$\text{Cr}_2\text{O}_7^{2-}$	dichromate			
	NO_3^-	nitrate	NO_2^- nitrite		N^{3-} nitride
MnO_4^- permanganate					
ClO_4^- perchlorate	ClO_3^-	chlorate	ClO_2^- chlorite	ClO^- hypochlorite	Cl^- chloride
BrO_4^- perbromate	BrO_3^-	bromate	BrO_2^- bromite	BrO^- hypobromite	Br^- bromide
IO_4^- periodate	IO_3^-	iodate	IO_2^- iodite	IO^- hypoiodite	I^- iodide
$\text{C}_2\text{H}_3\text{O}_2^-$ (or CH_3COO^-)		acetate		SCN^-	thiocyanate
HCO_3^-		hydrogen carbonate (bicarbonate)		CN^-	cyanide
HSO_4^-		hydrogen sulfate (bisulfate)		NH_4^+ (note positive charge)	ammonium
OH^-		hydroxide			

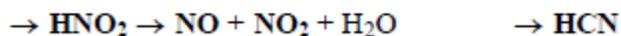
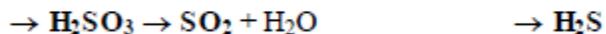
Per__ate: has one more oxygen than "ATE"

__ate: most common form (MCF)

__ite: one less oxygen than the "ate"

Formula	Name	Ion	Ion Name
H ₂ SO ₄	sulfuric acid	SO ₄ ²⁻	sulfate ion
H ₂ SO ₃	sulfurous acid	SO ₃ ²⁻	sulfite ion
HNO ₃	nitric acid	NO ₃ ¹⁻	nitrate ion
HNO ₂	nitrous acid	NO ₂ ¹⁻	nitrite ion
H ₃ PO ₄	phosphoric acid	PO ₄ ³⁻	phosphate ion
H ₂ CO ₃	carbonic acid	CO ₃ ²⁻	carbonate ion
HMnO ₄	permanganic acid	MnO ₄ ¹⁻	permanganate ion
HCN	hydrocyanic acid	CN ¹⁻	cyanide ion
HOCN	cyanic acid	OCN ¹⁻	cyanate ion
HSCN	thiocyanic acid	SCN ¹⁻	thiocyanate ion
HC ₂ H ₃ O ₂	acetic acid	C ₂ H ₃ O ₂ ¹⁻	acetate ion
H ₂ C ₂ O ₄	oxalic acid	C ₂ O ₄ ²⁻	oxalate ion
H ₂ CrO ₄	chromic acid	CrO ₄ ²⁻	chromate ion
H ₂ Cr ₂ O ₇	dichromic acid	Cr ₂ O ₇ ²⁻	dichromate ion
H ₂ S ₂ O ₃	thiosulfuric acid	S ₂ O ₃ ²⁻	thiosulfate ion
H ₃ AsO ₄	arsenic acid	AsO ₄ ³⁻	arsenate ion
H ₃ AsO ₃	arsenous acid	AsO ₃ ³⁻	arsenite ion

GASES THAT FORM



**** Make sure you know that ANY “carbonate” breaks down into CO₂ and H₂O or formed from those two things!****

Solubility Rules are MUST KNOW!!! You CAN'T do AP without these.

Solubility Rules—Mostly Soluble

Element/Ion- Compounds with these ions are soluble.	Exceptions- The ion STOPS being soluble when partnered with these ions.
Group I, NH_4^+ NO_3^- CH_3COO^- ($\text{C}_2\text{H}_3\text{O}_2^-$)	None (Takes precedence)
Cl^- , Br^- , I^-	Pb^{2+} Ag^+ Hg_2^{2+} “heavy metal bad guys”
SO_4^{2-}	Pb^{2+} Ag^+ Hg_2^{2+} Ca^{2+} Ba^{2+} Sr^{2+}
F^-	Pb^{2+} Ag^+ Hg_2^{2+} [Ca^{2+} Ba^{2+} Sr^{2+} Be^{2+} Mg^{2+}] <i>group II</i>

ALL OF THESE ARE 100% SOLUBLE!!

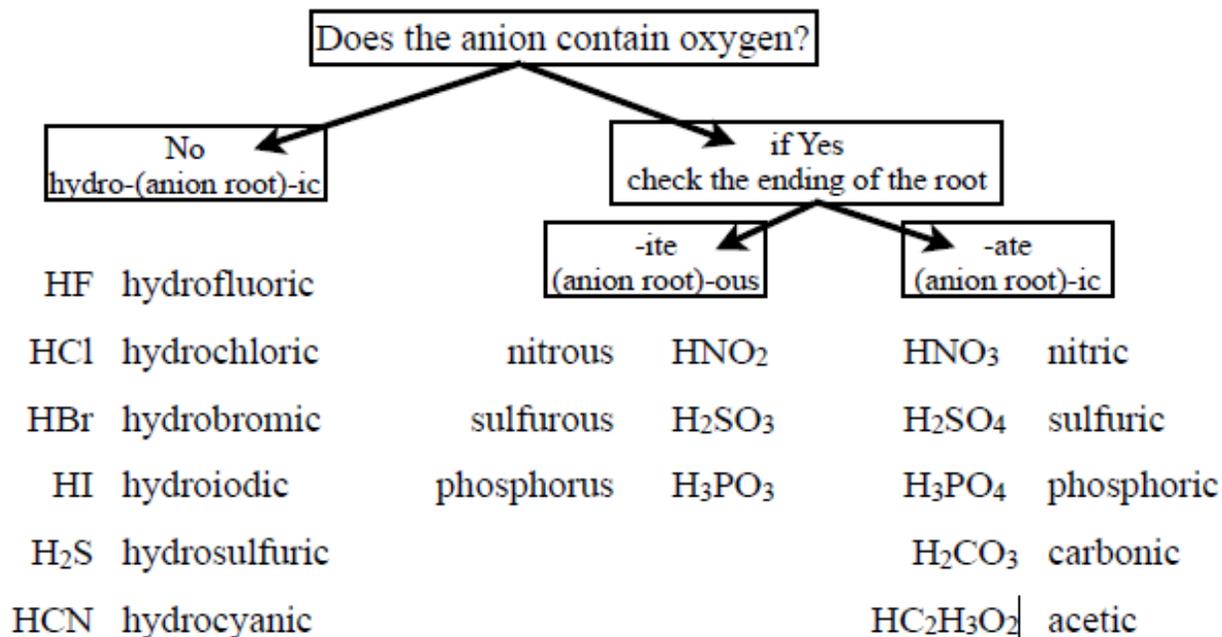
<u>Strong Acids</u>	<u>Strong Bases</u>	<u>Stinky gases</u>
HCl, HBr, HI	Group 1 hydroxides	*** $\text{H}_2\text{CO}_3 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ ***
Sulfuric Nitric	Calcium, Strontium	$\text{NH}_4\text{OH} \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
Perchloric	Barium hydroxides	$\text{H}_2\text{SO}_3 \rightarrow \text{H}_2\text{O} + \text{SO}_2$

Practice Section #1: Determine if the following will be soluble (aq) or insoluble (s).

- | | |
|----------------------------|------------------------------|
| 1. Calcium chloride _____ | 7. Barium Sulfate _____ |
| 2. Magnesium bromide _____ | 8. Galium Oxide _____ |
| 3. Barium nitrate _____ | 9. Lithium fluoride _____ |
| 4. Sodium carbonate _____ | 10. Lead(IV) Fluoride _____ |
| 5. Barium phosphate _____ | 11. Cesium acetate _____ |
| 6. Silver chloride _____ | 12. Aluminum hydroxide _____ |

Naming Acids will take you a LONG ways when we get to Ch. 14 and 15. There will be a quiz Week 4 of Naming Acids.

Demystifying the Naming of Acids (refer to page 61–62 in text)



Oxyhalogen Acids			
Formula	Oxy name	Ion	Ion name
HClO	hypochlorous acid	ClO ¹⁻	hypochlorite ion
HClO ₂	chlorous acid	ClO ₂ ¹⁻	chlorite ion
HClO ₃	chloric acid	ClO ₃ ¹⁻	chlorate ion
HClO ₄	perchloric acid	ClO ₄ ¹⁻	perchlorate ion

Br, I, can be substituted for Cl. F may form hypofluorous acid and the hypofluorite ion.

The Seven Strong Acids

memorize them

HNO ₃ nitric acid	HCl hydrochloric acid
H ₂ SO ₄ sulfuric acid	HBr hydrobromic acid
HClO ₃ chloric acid	HI hydroiodic acid
HClO ₄ perchloric acid	

III. Practice with acids! Remember:

-IC from -ATE	-OUS from -ITE	HYDRO-, -IC from -IDE
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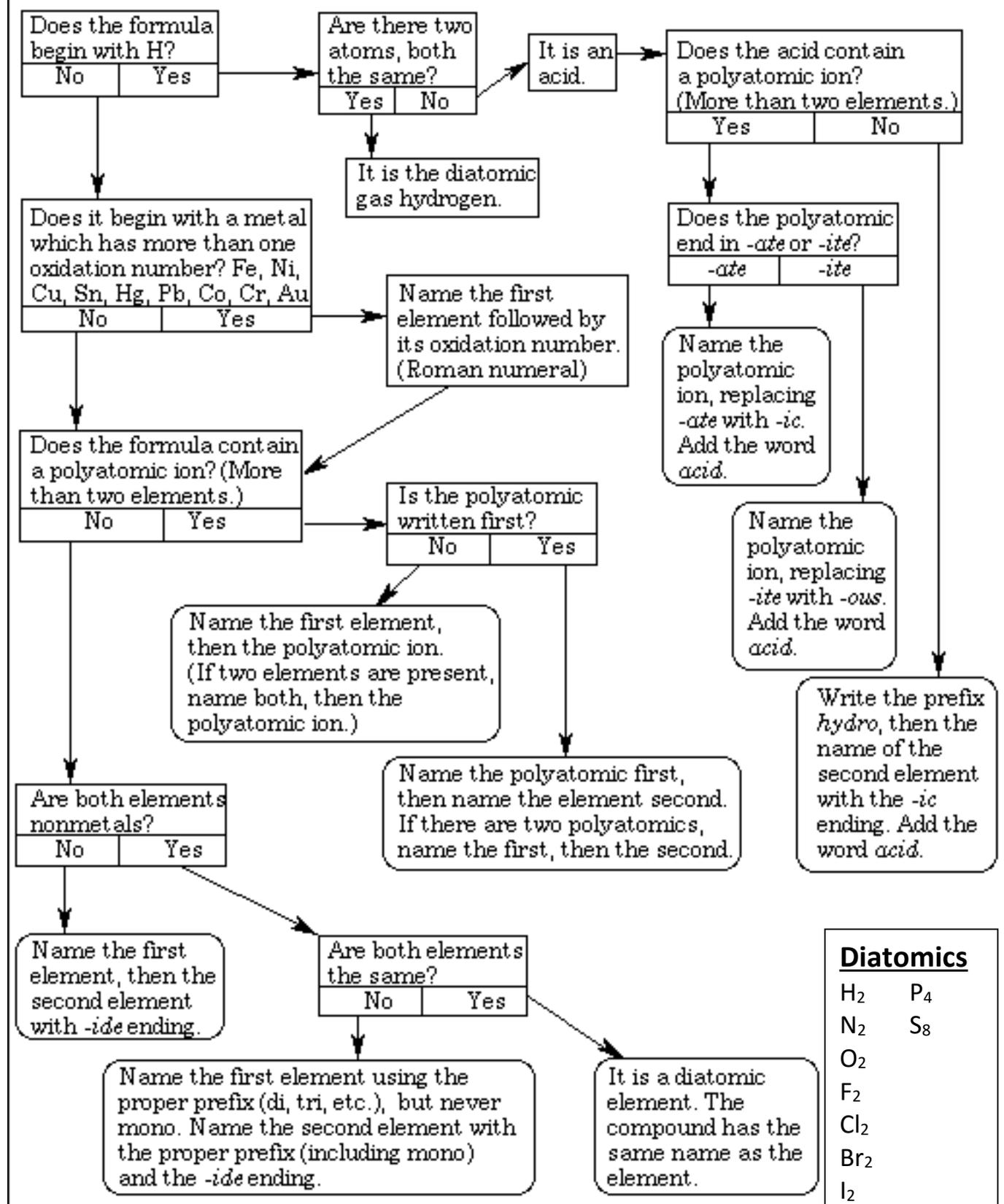
Practice Section #2: Give the name or formula for the following acids.

1. H_2S _____
2. HBr _____
3. HClO_2 _____
4. H_3PO_4 _____
5. HF _____
6. $\text{HC}_2\text{H}_3\text{O}_2$ _____
7. Chlorous acid _____
8. Hydrosulfuric acid _____
9. Nitrous acid _____
10. Hypochlorous acid _____
11. Hydronitric acid _____
12. Perchloric acid _____
13. Hydrocyanic acid _____
14. Phosphoric acid _____
15. Hydrophosphoric acid _____

Naming Compounds is one of the most important things!!

Flow Chart for Naming Simple Inorganic Compounds

The flowchart is adapted from p. 131-132 of the February 1983 issue of the *Journal of Chemical Education*.



Diatomics

H ₂	P ₄
N ₂	S ₈
O ₂	
F ₂	
Cl ₂	
Br ₂	
I ₂	

Practice Section #3: Give the name for the following simple organic compounds In the first column, classify each of the following as molecular (M) or ionic (I). In the second column, name each compound.

	M or I	Name		M or I	Name
1) CaF ₂			10) SrI ₂		
2) P ₄ O ₁₀			11) CO		
3) K ₂ S			12) Cs ₂ Po		
4) NaH			13) ZnAt ₂		
5) Al ₂ Se ₃			14) P ₂ S ₃		
6) N ₂ O			15) AgCl		
7) O ₂ F			16) Na ₃ N		
8) SBr ₆			17) Mg ₃ P ₂		
9) Li ₂ Te			18) XeF ₆		
1) FeSO ₃			16) Fe ₂ O ₃		
2) LiNO ₃			17) (NH ₄) ₂ SO ₃		
3) SrCl ₂			18) Ca(MnO ₄) ₂		
4) AgBr			19) PF ₅		
5) KClO ₃			20) LiH		
6) MgCO ₃			21) KClO ₃		
7) BaO ₂			22) NaBrO ₂		
8) KO ₂			23) Ca ₃ (PO ₄) ₂		
9) SnO ₂			24) LiClO ₄		
10) Ni ₃ (PO ₄) ₂			25) Fe(IO ₂) ₃		
11) Pb(OH) ₂			26) N ₂ I ₅		
15) S ₈			30) NH ₄ BrO ₃		

Basic Types of Chemical Reactions:

1. Synthesis (Composition):

- two or more elements or compounds may combine to form a more complex compound.
- **Basic form: $A + X \rightarrow AX$**

Examples of synthesis reactions:

1. Metal + oxygen \rightarrow metal oxide [Ionic Compound]
EX. $2\text{Mg}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{MgO}_{(s)}$
2. Nonmetal + oxygen \rightarrow nonmetallic oxide [Covalent Molecule]
EX. $\text{C}_{(s)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)}$
3. Metal + sulfur \rightarrow metal sulfide [Ionic Compound]
EX. $8\text{Ba}_{(s)} + \text{S}_{8(g)} \rightarrow 8\text{BaS}_{(s)}$
4. Metal + halogen \rightarrow Metal halide
EX. - Group 1: $2\text{Li}_{(s)} + \text{F}_{2(g)} \rightarrow 2\text{LiF}_{(s)}$
EX. - Group 2: $\text{Mg}_{(s)} + \text{I}_{2(g)} \rightarrow \text{MgI}_{2(s)}$
5. **Metal oxide + water \rightarrow metallic hydroxide**
EX. $\text{MgO}_{(s)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{Mg}(\text{OH})_{2(s)}$
6. **Nonmetallic oxide + water \rightarrow oxyacids**
EX. $\text{CO}_{2(g)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{H}_2\text{CO}_{3(aq)}$
7. Cation from an Acid (Metal) + Anion from a Base (nonmetal) \rightarrow salt
EX. $2\text{Na}_{(s)} + \text{Cl}_{2(g)} \rightarrow 2\text{NaCl}_{(s)}$
8. A few nonmetals combine with each other.
EX. $2\text{P}_{(s)} + 3\text{Cl}_{2(g)} \rightarrow 2\text{PCl}_{3(g)}$

These two reactions must be remembered:

1. $\text{N}_{2(g)} + 3\text{H}_{2(g)} \rightarrow 2\text{NH}_{3(g)}$ (Haber-Bosch Process)
2. $\text{NH}_{3(g)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{NH}_4\text{OH}_{(aq)}$

2. Decomposition:

- A single compound breaks down into its component parts or simpler compounds.
- **Basic form: $AX \rightarrow A + X$**

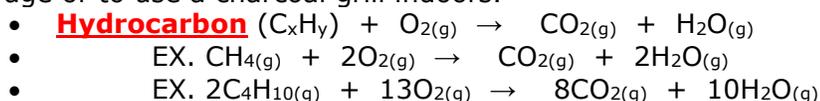
Examples of decomposition reactions:

1. **Metallic carbonates, when heated, form metallic oxides and $\text{CO}_{2(g)}$.**
EX. $\text{CaCO}_{3(s)} \rightarrow \text{CaO}_{(s)} + \text{CO}_{2(g)}$
2. **Most metallic hydroxides, when heated, decompose into metallic oxides and water. EXCEPT Group 1 Hydroxides!**
EX. $\text{Ca}(\text{OH})_{2(s)} \rightarrow \text{CaO}_{(s)} + \text{H}_2\text{O}_{(g)}$
3. **Metallic chlorates, when heated, decompose into metallic chlorides and oxygen gas.**
EX. $2\text{KClO}_{3(s)} \rightarrow 2\text{KCl}_{(s)} + 3\text{O}_{2(g)}$
4. **Some acids, when heated, decompose into nonmetallic oxides and water.**
EX. $\text{H}_2\text{SO}_4 \rightarrow \text{H}_2\text{O}_{(l)} + \text{SO}_{3(g)}$
5. **Some oxides, when heated, decompose to the elemental metal and O_2 .**
EX. $2\text{HgO}_{(s)} \rightarrow 2\text{Hg}_{(l)} + \text{O}_{2(g)}$
6. Some decomposition reactions are produced by **electricity**. This is called **electrolysis**
EX. $2\text{H}_2\text{O}_{(l)} \rightarrow 2\text{H}_{2(g)} + \text{O}_{2(g)}$
EX. $2\text{NaCl}_{(l)} \rightarrow 2\text{Na}_{(s)} + \text{Cl}_{2(g)}$

: Use the **solubility rules** to decide whether a product of an ionic reaction is insoluble in water and will thus form a precipitate. If a compound is soluble in water then it should be shown as being in aqueous solution, or left as separate ions. It is, in fact, often more desirable to show only those ions that are actually taking part in the actual reaction. Equations of this type are called **net ionic equations**. (This is a topic that will be covered to a much greater extent in AP Chemistry)

5. Combustion of Hydrocarbons:

- Another important type of reaction, in addition to the four types above, is that of the combustion of a hydrocarbon. When a hydrocarbon is burned with **sufficient oxygen supply**, the products are **always carbon dioxide and water vapor**. If the supply of oxygen is low or restricted, then carbon monoxide will be produced. This is why it is so dangerous to have an automobile engine running inside a closed garage or to use a charcoal grill indoors.



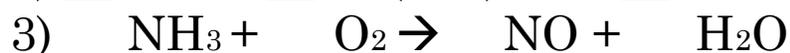
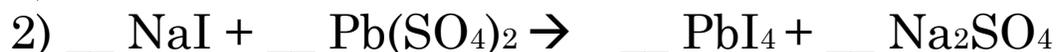
NOTE:

- **Tips to Balance**
 - Put a 2 in front of the Hydrocarbon
 - Balance as needed
 - **Reduce if you can**
- **Complete combustion** means the higher oxidation number is attained.
- **Incomplete combustion** means the lower oxidation number is attained.
- **The phrase "To burn"** means to add oxygen unless told otherwise

6. Acid/Base Neutralization reactions:

- This occurs when **one compound acquires H^+ from another** in the presence of an aqueous solution. The key is to **RECOGNIZE THE ACIDS! They will always be in aqueous solutions.**
 - Features
 - usually double displacement reactions (but not always)
 - one reactant must be an acid (have detachable H^+ ions)
 - one reactant must be a base (accept H^+ taken from the acid)
 - water is a product if the base is a hydroxide
 - Synonyms: acid-base reaction, proton transfer reaction
 - Examples
 - $H_2SO_4(aq) + 2 NaOH(aq) \rightarrow Na_2SO_4(aq) + 2 H_2O(l)$
 - $KOH + HC_2H_3O_2(aq) \rightarrow H_2O(l) + KC_2H_3O_2(aq)$
 - $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$

Practice Section #4: Balance the following equations by adding coefficients as needed. Some equations may already be balanced.



- 4) $__ \text{HNO}_3 + __ \text{Mg}(\text{OH})_2 \rightarrow __ \text{H}_2\text{O} + __ \text{Mg}(\text{NO}_3)_2$
- 5) $__ \text{H}_3\text{PO}_4 + __ \text{NaBr} \rightarrow __ \text{HBr} + __ \text{Na}_3\text{PO}_4$
- 6) $__ \text{CaO} + __ \text{MnI}_4 \rightarrow __ \text{MnO}_2 + __ \text{CaI}_2$
- 7) $__ \text{C}_2\text{H}_2 + __ \text{H}_2 \rightarrow __ \text{C}_2\text{H}_6$
- 8) $__ \text{VF}_5 + __ \text{HI} \rightarrow __ \text{V}_2\text{I}_{10} + __ \text{HF}$
- 9) $__ \text{OsO}_4 + __ \text{PtCl}_4 \rightarrow __ \text{PtO}_2 + __ \text{OsCl}_8$
- 10) $__ \text{Hg}_2\text{I}_2 + __ \text{O}_2 \rightarrow __ \text{Hg}_2\text{O} + __ \text{I}_2$

Practice Section #5: – Reaction Prediction Practice

I. Predict the products, write the formula equation, give the states of matter and then balance. Note the roman numerals following a metal designate the charge of the atom.

COMBUSTION

1. $\text{C}_4\text{H}_{10} + \text{oxygen} \rightarrow$
2. $\text{C}_7\text{H}_{14} + \text{oxygen} \rightarrow$

SYNTHESIS

1. sodium oxide + water \rightarrow
2. calcium + nitrogen \rightarrow
3. sulfur dioxide + water \rightarrow

DECOMPOSITION

1. Strontium carbonate \rightarrow
2. Mercury (II) oxide \rightarrow
3. aluminum chlorate \rightarrow

DOUBLE REPLACEMENT/ACID_BASE

1. Aluminum sulfate + calcium hydroxide \rightarrow

2. Sodium hydroxide + sulfuric acid →
3. sodium sulfide + manganese (VI) acetate →
4. chromium(III) bromide + sodium sulfite →
5. barium hydroxide + hydrochloric acid →

SINGLE REPLACEMENT

1. Nickel + steam →
2. Chlorine + aluminum iodide →
3. Potassium + water →
4. Zinc + hydrochloric acid →

Trends in the Periodic Table

★ VERY IMPORTANT!!! ★

Justifying all of the trends on the periodic table can be simplified using these two generalizations:

- 1) Use Z_{eff} to justify trends across a period.
- 2) Use increased distance (greater value of n) to justify trends down a group.



of e- = more orbitals = more polarizable!!

Atomic radius refers to the distance between the nucleus and the outer edge of the electron cloud. It is influenced by the nuclear pull and the number of energy levels.

Atomic radii decrease as atomic numbers increase in any given period	
DO	DON'T
Teach students that the effective nuclear charge, Z_{eff} , increases the attraction of the nucleus and therefore pulls the electron cloud closer to the nucleus resulting in a smaller atomic radius.	Don't let students get away with simply stating that atomic radii decrease from left to right across a period.

Atomic radii increase as atomic number increases down a column or group	
DO	DON'T
Teach students that the increased number of energy levels (n) increases the distance over which the nucleus must pull and therefore reduces the attraction for electrons.	Don't let students get away with simply saying that radii increase down a column.
Teach students that full energy levels provide some shielding between the nucleus and valence electrons.	Don't let students use shielding for explanations across a period. Only full energy levels, not full sublevels, are of concern in a shielding argument.

Ionization energy refers to the energy needed to remove an electron from a *gaseous* atom or ion, i.e. an isolated one, not part of a solid, liquid or a molecule. It is *always* endothermic.

Ionization energy increases as atomic number increases in any given period	
DO	DON'T
Teach students that the effective nuclear charge, Z_{eff} , increases the attraction of the nucleus and therefore holds the electrons more tightly.	Don't let them get away with simply stating that ionization energy increases from left to right across a period.
Teach students the exceptions that occur between groups II and III and V and VI.	Don't let them think that the trend is unwavering.
1) A drop in IE occurs between groups II and III because the p electrons do not penetrate the nuclear region as greatly as s electrons do and are therefore not as tightly held.	1) Don't let them state that p electrons are farther away from the nucleus.
2) A drop in IE occurs between groups V and VI because the increased repulsion created by the first pairing of electrons outweighs the increase in Z_{eff} and thus less energy is required to remove the electron.	2) Don't let them state that the atoms in group V are more stable because they have a half filled sublevel. This is wrong, wrong, wrong!

Ionization energy decreases as atomic number increases down a column or group

DO	DON'T
Teach students that the increased number of energy levels (n) increases the distance over which the nucleus must pull and therefore reduces the attraction for electrons.	Don't let students get away with simply saying that IE decreases down a column.
Teach students that full energy levels provide some shielding between the nucleus and valence electrons.	Don't let students use shielding for explanations across a period. Only full energy levels, not full sublevels, are of concern in a shielding argument.

Electron affinity is NOT the opposite of ionization energy, but involves the addition of an electron to a gaseous atom or ion, which can be exothermic or endothermic. The exothermic values can be confusing for students since -500 kJ represents a higher electron affinity than -100 kJ. You can tell students to consider the absolute value of the energy term since the negative sign is simply indicating the direction of energy flow.

Electronegativity is a property (there are several scales) which measures the attraction of an atom for the *pair* of outer shell electrons in a covalent bond with another atom. Electronegativity patterns are the same as electron affinity patterns for the same reasons. Both of these properties focus on the attraction that the nucleus has for electrons.

Electronegativity increases as atomic numbers increase in any given period

DO	DON'T
Teach students that the effective nuclear charge, Z_{eff} , increases the attraction of the nucleus and therefore it strengthens the attraction for the electrons.	Don't let them get away with simply stating that electronegativity increases from left to right across a period.

Electronegativity decreases as atomic number increases down a column or group

DO	DON'T
Teach students that the increased number of energy levels (n) increases the distance over which the nucleus must pull and therefore reduces the attraction for electrons.	Don't let students get away with simply saying that electronegativity decreases down a column.
Teach students that full energy levels provide some shielding between the nucleus and valence electrons.	Don't let students use shielding for explanations across a period. Only full energy levels, not full sublevels, are of concern in a shielding argument.

Final thoughts

Students often have trouble immediately recognizing the difference between the two species given. Teach them to follow these three steps EVERY time they start one of these questions and their scores are sure to go up.

- 1) Locate both elements on the periodic table and note the energy level (n) and sublevel of their valence electrons.
- 2) Do they have the same or different n values?
- 3) If same n : argue with Z_{eff} ; if different n : argue with n vs. n distances.

MUST KNOW!!!

Groups of the periodic table: For each of the following groups, know their location, common uses, and common properties.

Group	Location	Common Uses	Common Properties
Alkali Metals			
Alkaline Earth Metals			
Transition Metals			
Halogens			
Noble Gases			
Metals			
Non-metals			

Practice Section #6: – Periodic Trends Place these elements (S, Se, I, Ca and Be.) in order of the following.

Please note it will NOT be enough to memorize these rules you will have to EXPLAIN them, BUT we will spend a great deal of time on the explanation. It is important that you already understand the trends:

- A. increasing atomic radius
- B. decreasing ionization energy
- C. increasing electronegativity

Review and know the following 7 shapes and their angles.

Bonding Electron Pairs	Lone Pairs	Electron Domains	Shape	Ideal Bond Angle (example's bond angle)	Example	Image
2	0	2	linear	180°	BeCl ₂	
3	0	3	trigonal planar	120°	BF ₃	
4	0	4	tetrahedral	109.5°	CH ₄	
3	1	4	trigonal pyramidal	109.5° (107.5°)	NH ₃	
2	2	4	bent	109.5° (104.5°)	H ₂ O	
5	0	5	trigonal bipyramidal	90°, 120°	PCl ₅	
6	0	6	octahedral	90°	SF ₆	

**There are more shapes to know but they are all based on these: memorize

them!**

Lewis Structure Steps

1. List each element in the formula and how many of each are present
2. Determine the number of valence electrons for each element (valence # is group #)
3. Determine the **TOTAL number of electrons** for that compound
4. Draw the skeleton structure
 - a. Determine the central atom (the atom with the most available bonds and/or the **least** electronegative)
 - b. Connect all the other elements to the central atom with a line
 - i. Subtract the number of bonded electrons from the total number
 - One line = two electrons
 - c. Give all the **outer** atoms a full octet (except Hydrogen)
 - i. Subtract the number of octet electrons from the total number
 - d. Put all of the extra electrons on the central atom
 - i. Do this until you're out of electrons
5. Determine the shape of the molecule and **draw the correct representation** based on the shape you determined.

**Notes:

- If there is not an octet on the central atom or outer electrons, look for places to make double bonds, and even triple bonds in some cases (rare)
- Exceptions to the Octet Rule:
 - Hydrogen and Helium only need **2 electrons**
 - Boron only has to have **6 electrons**
 - Phosphorus, Sulfur, Krypton and Xenon can have **MORE than 8 (expanded octet)**

Polar vs Nonpolar:

A polar molecule has an uneven charge distribution (the electrons are not shared equally) within the molecule. This creates a partially negative portion of the molecule as well as a partially positive end.

A nonpolar molecule has an even charge distribution (the molecules are shared equally) within the molecule so there are no charged areas within the atom. It is necessary to look at the Lewis structure to determine polarity. Typically

molecules that have no lone pairs and the same atoms around the central atom are considered to be polar because these molecules show a symmetrical arrangement

Intermolecular Forces:

These are the attractive forces BETWEEN MOLECULES. These are not as strong as actual bonds but they do allow for groups of molecules to be held together such as when forming a solid. A group of molecules may participate in more than one type of intermolecular force but for now we will differentiate the following by these three categories:

1. Hydrogen Bonds: These are used on polar molecules that contain a Hydrogen AND either a Fluorine, oxygen or nitrogen. Polar molecules contain a slight charge, so these opposite charges are attracted to each other.
2. Dipole-Dipole Forces: These are used on all other polar molecules. These work much the same as hydrogen bonds but they are not as strong.
3. London Dispersion Forces: These are used by non-polar molecules. These occur due to an instantaneous dipole (momentary polarity) and the attractive forces that occur between two oppositely charged particles.

Practice Section #7: Complete the following table giving the correct Lewis structure, shape, bond angle, indicating polar or nonpolar, and which intermolecular force would hold be present.

Name of compound	Lewis Structure	Shape and Bond angle(s)	Polar, non-polar, or N/A	Type of IMF
Cl ₄				
CO ₂				
HCN				
CH ₂ O				
NO ₃ ⁻¹				

NH ₃				
PCl ₅				
O ₂				

Practice Section #8: Electron Configuration (this is 1s², 2s², 2p⁶, 3s²...& so on)

- For Se write:

A. the complete electron configuration



B. the noble gas electron configuration [Pick the NG PREVIOUS TO THE ELEMENT]

C. the orbital diagram from the noble gas electron configuration (this one is the arrows: Up and down)

D. the dot diagram

- For the following elements, draw the Lewis dot structures.

A. N



B. F

C. Ca

D. He

Significant figure rules

- Exact numbers: Any number that has been counted, so there is no uncertainty.

1. 35 gummy bears, 24 drops,
- Measured numbers: The number has an uncertainty based on the precision of the equipment. This means that the last number is estimated.
 1. 1.45 ± 0.01 cm, or 2,360 mL or 0.0055 atm
 2. A number is significant (or measured) if...
 - it is a number other than zero
 - It is a zero between two significant numbers
 - It is a zero after the decimal and after significant numbers
 - A number is not significant (or measured) if it acts only as a placeholder..
 - It is after significant numbers but before the decimal
 - It is after the decimal but before significant figures
- Mnemonic: decima**L**-if the number has a decimal begin counting sig figs with the first non-zero from the **Left**. If the number has no decimal begin counting with the first non-zero from the right.

Practice Section #9: Significant Figures

1. Give the number of sig figs in each of the following numbers

a. 123 _____	f. 0.009 _____
b. 0.078 _____	g. 23,000. _____
c. 89007 _____	h. 34,000 _____
d. 12,000 _____	i. 34.89 _____
e. 1,000,000,000.0 _____	j. 101 _____
2. Do the following calculations giving the answer in the appropriate number of sig figs.

a. $1.23 + 75 =$ _____	e. $0.887 + 0.3 =$ _____
b. $1.89 - .20 =$ _____	f. $2340 - 100 =$ _____
c. $45.6 \times 8.2 =$ _____	g. $12.45 \times 3 =$ _____
d. $234 / 0.29 =$ _____	h. $25,600 / 3.0 =$ _____
3. Do the following calculations giving the answer in the appropriate number of sig figs
 - a. $45.0 \times 9.0 + 89.22 / 75 =$ _____
 - b. $(2.88 + .5) \times (23,000 - 0.11) =$ _____
 - c. $0.8897 \times 2.15 + 0.002 / .1 =$ _____
 - d. $(8 + 9) / (34.0 - 20.) =$ _____

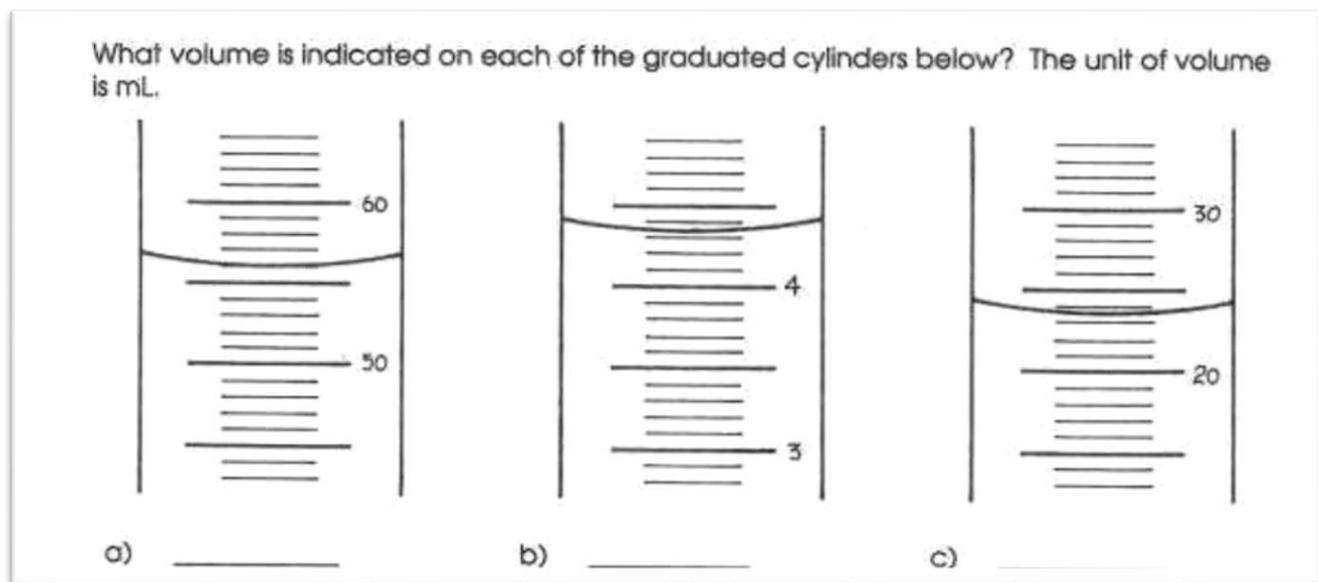
Practice Section #10: Directions: Solve each of the following problems. Show all work and write your answer using the correct number of significant figures and units.

- Convert 3.48×10^{20} molecules of SO_2 to moles. What is the mass of this

quantity?

- Calculate the following for quantities for 4.68g of $\text{Ca}_3(\text{PO}_4)_2$:
 - a. moles
 - b. Ca^{2+} ions (Use subscripts as Mole Ratios)

Practice Section #11: Practicing Measuring Liquid Volume. Give the volumes of the following graduated cylinders using the appropriate number of significant figures.



Pra **ESTIMATE ONE DIGIT PAST PRECISE KNOWN MEASUREMENT!!**
M **BE CAREFUL ON HOW THE MEASUREMENTS READ!!**

- V** - V must be in liters (change if necessary)
- Use **M** or mol/L as unit for molarity

- How many moles of solute are present in 100. ml of 1.50 M MgSO_4 ?
- What is the molarity of 35 g of iron (II) acetate dissolved in enough water to make 250 ml of solution?
- How many moles of LiF would be required to produce a 2.5 M solution with a volume of 1.5 L?

- How many moles of $\text{Sr}(\text{NO}_3)_2$ would be used in the preparation of 2.50 L of a 3.5 M solution?
- What is the molarity of a 500-ml solution containing 249 g of KI?
- How many grams of CaCl_2 would be required to produce a 3.5 M solution with a volume of 2.0 L?
- How many liters of solution are needed to make a 1.66 M solution containing 2.11 moles of KMnO_4 ?

Ⓢ Limiting Reagent (Reactant) - (LR)

- ◆ Controls the amt. of product formed
- ◆ Completely consumed in the rxn.
- ◆ “runs out” first
- ◆ Example:
People on plane → 300 people; 250 seats
Seats are the limiting factor

Ⓢ Limiting Reagent Problems

- ◆ Do mass – mass (g to g) calc. for all reactants.
- ◆ Whichever reactant produces the least is the Limiting Reagent (LR)

Ⓢ Excess Reactant (ER)

To find the amount of excess reactant leftover after a rxn:

- ◆ Do two mass-mass (g to g) problems to find LR
- ◆ Use LR to calculate excess reactant used.
- ◆ Subtract excess reactant used from original amt. of excess reactant = leftover excess reactant

Ⓢ Percent Yield

$$\text{◆ } \% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

◆ theoretical yield:

- maximum amt. of product (what you should “have gotten”)
- from mass-mass (g to g) problem

◆ actual yield:

- actual amt. of product (what you “got”)
- from lab result; or given in a problem

Practice Section #13: Solve each of the following problems. Show all work and write your answer using the correct number of significant figures and units. Write the equation if it is not given.

1. How many grams of potassium nitride are needed to make 500 mL of a 0.35 M solution?
2. Suppose that an excess of propane, C_3H_8 burns in 320 g of O_2 . How many moles of H_2O will be formed? $\underline{\hspace{1cm}} C_3H_8 + \underline{\hspace{1cm}} O_2 \rightarrow \underline{\hspace{1cm}} CO_2 + \underline{\hspace{1cm}} H_2O$
3. How many grams of sodium sulfate will be formed if you start with 200 grams of sodium hydroxide and you combine it with an excess of sulfuric acid?
4. What is the actual amount of magnesium chlorate that will be formed from the reaction of 5 moles of magnesium iodide and 25 g of calcium chlorate?
5. When I carry in a 5-gallon container of water from my car, I always wonder its weight. I looked up on the web and found that 1 lb = 0.453542 kg and 1 qt = .946353 Liter. Calculate its weight in pounds.
6. How many kilometers per minutes are in 550 meters per second?

7. Suppose a solution containing 4.50g of sodium phosphate is mixed with a solution containing 3.75g of barium nitrate. How many grams of barium phosphate can be produced? ****Write the reaction and solve for the LR****. **Then convert the LR to g of barium phosphate****

8. Over the years, the thermite reaction (**mixing of solid iron (III) oxide with aluminum metal**) has been used for welding railroad rails, in incendiary bombs, and to ignite solid-fuel rocket motors.

A. Write a balanced equation representing the reaction. (It's a Single Replacement Rxn)

B. What masses of iron (III) oxide and aluminum must be used to produce 15.0 g of iron?

C. What is the maximum mass of aluminum oxide that could be produced?

9. **Given: 50 ml of 2.0 M nitric acid reacts with 15 g of aluminum hydroxide**

A. Write a balanced equation for the reaction.

B. Write the net ionic equation. ****Look up how to do or ask****

C. Which of the reactants would be the limiting reagent?

D. How many grams of water would be produced?

Other types of problems that will come back during the year:

Practice Section #14: Solve each of the following problems. Show all work and write your answer using the correct number of significant figures and units. Write the equation if it is not given.

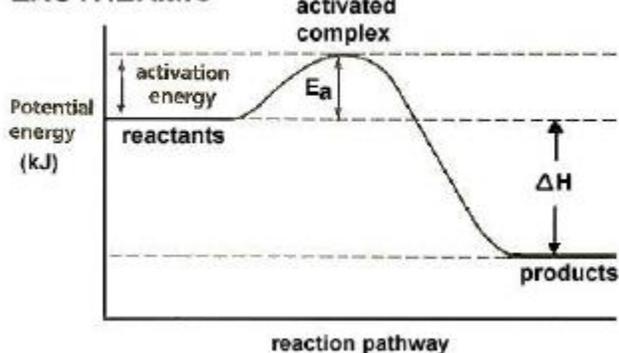
1. A sample of diborane gas (B_2H_6) has a pressure of 345 torr at a temperature of $-15^\circ C$ and a volume of 3.48 L. If conditions are changed so that the temperature is $36^\circ C$ and the pressure is 268 torr, what will be the volume of the sample? ****Combined Gas Law $P_1V_1T_2 = P_2V_2T_1$... Remember to change the $^\circ C$ to Kelvin ($K = C + 273$)**

2. The specific heat capacity of graphite is $0.71 J/^\circ C \cdot g$. Calculate the energy (in calories) required to raise the temperature of 1800 g of graphite by $100.0^\circ C$. **** $q = m \times c_p \times \Delta T$ **... Solve for q ; $m = \text{mass}$**

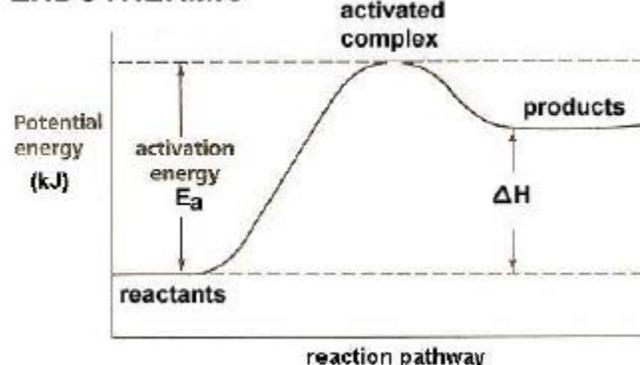
$C_p = \text{specific heat}$; and $\Delta T = \text{change in temp. (can leave in } ^\circ C)$

Thermochemistry Diagrams

EXOTHERMIC



ENDOTHERMIC



$$K_w = [\text{OH}^-][\text{H}^+] = 1 \times 10^{-14}$$

$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = -\log[\text{H}^+]$$

$$[\text{H}^+] = 10^{-\text{pH}}$$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

⊙ Pure H₂O has:

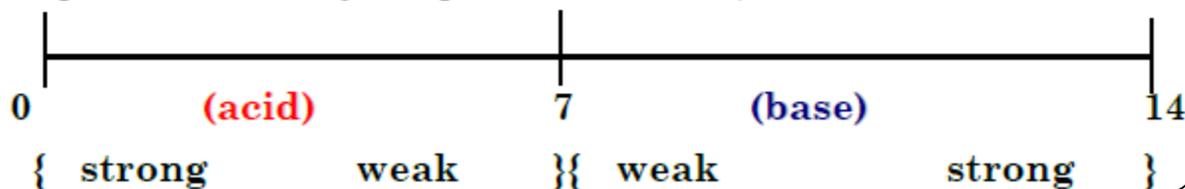
$$1 \times 10^{-7} \text{ M H}^+ \text{ (same as H}_3\text{O}^+) \text{ and } 1 \times 10^{-7} \text{ M OH}^-$$

⊙ Because $[\text{H}^+] = [\text{OH}^-]$ in water = neutral solution
(remember the brackets stand for concentration in molarity)

⊙ When $[\text{H}^+] > [\text{OH}^-]$; the solution is **acidic**

⊙ When $[\text{OH}^-] > [\text{H}^+]$; the solution is **basic**

⊙ pH is an easier way to express concentration; remember:



Very
Very
Very
Important

Acids, Bases, pH, and Titrations

3. What is the $[\text{H}^+]$, $[\text{OH}^-]$, pH, and pOH of a 0.005 M solution of lithium hydroxide?

Complete the following chart.

	[H ⁺]	pH	[OH ⁻]	pOH	Acidic or Basic
1.	10 ⁻⁵ M	5	10 ⁻⁹ M	9	Acidic
2.		7			
3.			10 ⁻⁴ M		
4.	10 ⁻² M				

Equilibrium & LeChatelier's Principle:

Explain any shift that would occur for the following and **explain why**:

Conditions:

- Concentration:** A. Increase/add → shifts away (to consume excess)
B. Decrease/Remove → shifts towards (to replace)
- Pressure:** A. Increase → shifts to side with less gaseous moles
B. Decrease → shifts to side with more gaseous moles
- Temperature:** – depends on if rxn is exothermic or endothermic

(Shift away from addition; Shift to Removal)

5. When Phosphorus pentachloride gas decomposes to form phosphorus trichloride gas and chlorine gas, 120 J of heat are released.

A. Write a balanced equation for this reaction.

B. more phosphorus pentachloride is added.

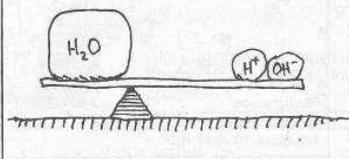
C. The temperature is decreased

D. The pressure is increased

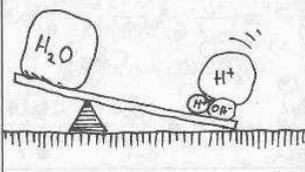
E. Chlorine gas is removed

Le Chatelier's Principle

YOU CAN THINK OF EQUILIBRIUM AS A BALANCED SEESAW WITH REACTANTS ON ONE SIDE AND PRODUCTS ON THE OTHER. IN THE LAST EXAMPLE, H_2O WAS ON THE LEFT, OH^- AND H^+ ON THE RIGHT.



IN THAT EXAMPLE, THE EQUILIBRIUM WAS DISTURBED BY ADDING H^+ TO THE RIGHT SIDE. WHAT HAPPENS THEN?



THE FRENCH CHEMIST HENRY LE CHATELIER HAS LEFT US A GENERAL PRINCIPLE FOR ANALYZING WHAT HAPPENS WHEN CHEMICAL EQUILIBRIUM IS DISTURBED.

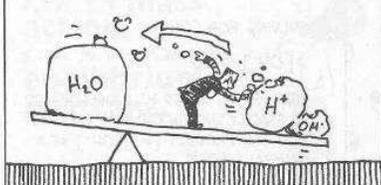
When an external stress is applied to a system at equilibrium, the process evolves in such a way as to reduce the stress.



FOR EXAMPLE, IF $aA + bB \rightleftharpoons cC + dD$ IS IN EQUILIBRIUM, THEN ADDING REACTANT A DRIVES THE REACTION TO THE RIGHT—CONSUMING MORE A.



IN OUR EXAMPLE, ADDING LOADS OF H^+ TO THE RIGHT-HAND SIDE OF $H_2O \rightleftharpoons H^+ + OH^-$ DROVE THE REACTION TO THE LEFT.



$[OH^-]$ FELL SHARPLY, AND EVERY OH^- ION THAT DISAPPEARED TOOK AN H^+ WITH IT, THEREBY LOWERING $[H^+]$.



Net Ionic Equations

Net ionic equations allow the spectator ions to be removed so as to visualize the reaction more effectively. Any compound that is aqueous (dissolved in water) would not remain as a bonded set, instead, the compound would dissolve into the appropriate ions. Any species that is an ion at the beginning and the end of the reaction should be removed. For each of the following, give the balanced equation, the full ionic equation (with the aqueous compounds written as ions) and the net ionic equation.



-
-
-



-
-
-

3. A small piece of calcium metal is added to hot distilled water
 - a.
 - b.
 - c.
4. A solution of copper(II) chloride is added to a solution of sodium sulfide.
 - a.
 - b.
 - c.
5. $C_4H_{10}O$ is burned in air.
 - a.
 - b.
 - c.
6. Chlorine gas is bubbled through a solution of potassium bromide.
 - a.
 - b.
 - c.
7. Solutions of strontium nitrate and sodium sulfate are mixed.
 - a.
 - b.
 - c.

Titrations

A titration allows us to determine the concentration of an unknown compound by comparing the unknown (analyte) to the standard solution (solution of known concentration, aka titrant). These can be solved using stoichiometry in which you compare the number of moles of the known concentration to the unknown via a mole ratio and then divide by the number of liters that were required to titrate. I recommend this version for AP chemistry because it is more universal and shows a better understanding of titrations. However, sometimes you can use the equation $MV_n = MV_n$. M : Molarity, V : Volume, and n : # of moles of ions.

For each of the following, calculate the unknown concentration.

1. What concentration of calcium hydroxide is present if 25.0 mL were titrated with 52.55 mL of 0.025 M hydrochloric acid?

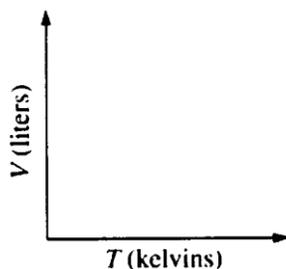
2. It requires 1.95 grams of sodium hydroxide to react with 75.00 mL sulfuric acid. What is the concentration of sulfuric acid? (Hint: Use stoichiometry. *Convert grams of sodium hydroxide to moles, then use the mole ratio of sodium hydroxide to sulfuric acid, then convert the number of moles of acid to molarity by dividing by the L of sulfuric acid*)
3. What is the concentration of 125.0 mL of nitric acid when 2.50 grams of barium hydroxide are needed to react?
4. How many milliliters of 0.75 M potassium hydroxide solution would you anticipate it takes to react with 50.0 mL of 2.50 M hydrobromic acid?

FUTURE THINGS TO COME IN AP CHEMISTRY!!

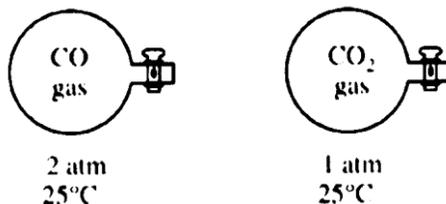
The following are examples of Free Response Type questions you would see on the AP test. These were selected because you have been exposed to the material and should be able to answer them. You are only required to attempt them, and remember we will be working all year to perfect your work on the FRQs. Clearly show the method used and the steps involved in arriving at your answers. It is to your advantage to do this, since you would be able to obtain partial credit if this were an actual exam. Attention should be paid to significant figures and all questions should have correct units with answers clearly labeled.

1. To increase the speed of a chemical reaction the temperature is often raised. Two factors are commonly cited as accounting for the increased rate of chemical reaction as the temperature is increased.
 - (a) List the two factors.
 - (b) Discuss the role of the two factors in increasing the rate of the chemical reaction.
 - (c) Which of the two is more important in increasing the rate of the chemical reaction? Explain your answer.

2. Answer the following questions about carbon monoxide, $\text{CO}(g)$, and carbon dioxide, $\text{CO}_2(g)$. Assume that both gases exhibit ideal behavior.
- Draw the complete Lewis structure (electron dot diagram) for the CO molecule and for the CO_2 molecule.
 - Identify the shape of the CO_2 molecule.
 - One of the two gases dissolves readily in water to form a solution with a pH below 7. Identify the gas and account for this observation by writing a chemical equation.
 - A 1.0 mol sample of $\text{CO}(g)$ is heated at constant pressure. On the graph below, sketch the expected plot of volume versus temperature as the gas is heated.



- (e) Samples of $\text{CO}(g)$ and $\text{CO}_2(g)$ are placed in 1 L containers at the conditions in the diagram below.



- Indicate whether the average kinetic energy of the CO_2 is greater than, equal to, or less than the average kinetic energy of the $\text{CO}(g)$ molecules.
- Indicate whether the number of $\text{CO}_2(g)$ molecules is greater than, equal, or less than the number of $\text{CO}(g)$ molecules. Justify your answer.

(Warning: I know this question is difficult. Try your best!)

3. Using a lab procedure, you are asked to find the concentration of magnesium hydroxide in a sample of Maylox[®]. You have the following equipment available to you during the lab.
Buret, graduated cylinders (10mL and 100 mL), 250 mL beaker, 250 mL Erlenmeyer flask, volumetric pipette, Ringstand, Bunsen burner, hot plate, distilled water, Sodium hydroxide, hydrochloric acid, phenolphthalein(indicator), and a mortar and pestle.
- Give the brief procedure needed to perform the titration safely during this lab.
 - Give the data you will need to record during the lab in order to determine the concentration of magnesium hydroxide in your sample of Maylox[®].
 - Assuming that you used 10.05 mL of 1.50 M HCl to react with 50.0 mL of Maylox. Give the concentration of magnesium hydroxide found in the sample of Maylox[®].
 - In the flask below, draw the ions that would remain the flask at the completion of the titration.
 - During the lab, the following errors occurred. State whether the concentration would be higher, lower, or unaffected and explain why.
 - The sample of Maylox spilled after weighing.
 - Not all of the Maylox dissolved in the hydrochloric acid.
 - Too much titrant was added to the analyte.

