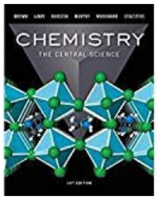
Course 3651
AP Chemistry
Required Summer Packet

Books Required



Brown, Theodore, et al. *Chemistry: The Central Science (14th Edition)* (without mastering Chemistry access),
ISBN-13: 9780134650951, Pearson , 2017.

Read Chapter 1 – Textbook
Read Chapter 2 – Textbook
Read Chapter 3 – Textbook

Pocket File (13 pocket)

Sections in Pocket File

Section 1- Chapter 1,2,3
Section 2- Redox Chapter 20.1-20.2 / Atomic Structure
Section 3- Aq Rxs and Solution Chem Chapter 4
Section 4- Bonding & Geometry & Forces Chapter 8 & 9 & 11
Section 5- Gases -Chapter 10
Section 6- Thermochemistry -Chapter 5 & Chapter 19
Section 7- Kinetics-Chapter 14
Section 8- Chemical Equilibrium –Chapter 15
Section 9- Equilibrium –Chapter 16
Section 10-Acids & Bases -Chapter 17
Section 11- Electrochemistry -Chapter 20
Section 12- Nuclear/Organic Chapter 21,22,24
Section 13-Reactions & Practice AP Exams

AP Chemistry Summer Review

Welcome to AP Chemistry. Summer review work is meant to get us off to a running start when fall arrives. You should already know most of the material in this packet from your first year chemistry class. Nonetheless, START NOW to make your way leisurely through this review. Do not think you should wait until the very end of the summer so it is "fresh in your mind." The LONGER it is in your mind, the better it will stick. Email me with any questions or concerns. I will try to respond ASAP.

Be prepared for an exam on this material the first week of school.

Nomenclature

You will be more successful in AP chemistry if you can name chemicals from their formulas, and if you can write chemical formulas from the name of the chemical. You don't need to know all polyatomic ions – just a short list will be helpful. You should be able to do this with only the assistance of a periodic table, after *memorizing* the polyatomic ions in the charts below.

Polyatomic Ions -

Memorize the shaded ions (and learn the pattern so you can easily memorize their companions)

By learning the four shaded "-ate" ions in the table to the right, **and** knowing that one less oxygen (same charge) turns the name to *–ite*, **and** two less oxygens (when possible) turns the name to *hypo*–xxx–ite **and** one more oxygen (when possible) turns the name to *per*–xxx–ate will make learning all eighteen ions in the chart below as easy as learning just four.

Polyatomic Ions to Memorize *(Use the pattern to help)*			
hypo- (2 less O)	**-ite (1 less O)**	**-ate**	**per- (1 more O)**
	nitrite NO_2^-	nitrate NO_3^-	
	sulfite SO_3^{2-}	sulfate SO_4^{2-}	
	phosphite PO_3^{3-}	phosphate PO_4^{3-}	
hypochlorite ClO^-	chlorite ClO_2^-	chlorate ClO_3^-	perchlorate ClO_4^-
hypobromite BrO^-	bromite BrO_2^-	bromate BrO_3^-	perbromate BrO_4^-
hypoiodite IO^-	iodite IO_2^-	iodate IO_3^-	periodate IO_4^-

Memorize the six extra polyatomic ions in the second table below, and don't forget ammonium in the table by itself.

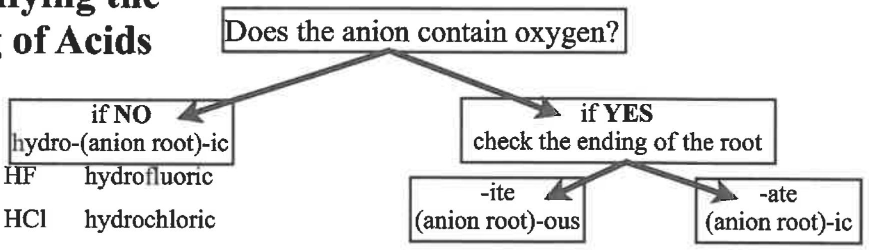
Odd Companions or No Companion	
hydroxide OH^-	
cyanide CN^-	
acetate $C_2H_3O_2^-$	
carbonate CO_3^{2-}	bicarbonate HCO_3^-
permanganate MnO_4^- *purple color*	

and don't forget
ammonium NH_4^+

Acids and Bases

You will need to be able to name acids. Knowing your polyatomic ions is critical in naming the acids. Use the chart below to review the pattern and method of naming.

Demystifying the Naming of Acids



HF	hydrofluoric
HCl	hydrochloric
HBr	hydrobromic
HI	hydroiodic
H_2S	hydrosulfuric
HCN	hydrocyanic

nitrous	HNO_2
sulfurous	H_2SO_3
phosphorus	H_3PO_3
chlorous	$HClO_2$
bromous	$HBrO_2$

nitric	HNO_3
sulfuric	H_2SO_4
phosphoric	H_3PO_4
carbonic	H_2CO_3
acetic	$HC_2H_3O_2$

You should memorize the seven strong acids in the table. The strong bases are group I and group II metal hydroxides. Hopefully you are aware that to be a strong acid or strong base means that when dissolved in water, the molecules are fully ionized. This means that the compound will dissociate completely into ions when in solution. This is important to recognize when writing net ion ionic equations.

Seven Strong Acids

memorize them
(assume all other acids are weak)

HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydroiodic acid
HNO_3	nitric acid
H_2SO_4	sulfuric acid
$HClO_3$	chloric acid
$HClO_4$	perchloric acid

Strong Bases

memorize them ***(Group I and II hydroxides)***

LiOH	lithium hydroxide	*Be & Mg hydroxides are not very useful, since they are not soluble*	
NaOH	sodium hydroxide		
KOH	potassium hydroxide	$Ca(OH)_2$	calcium hydroxide
RbOH	rubidium hydroxide	$Sr(OH)_2$	strontium hydroxide
CsOH	cesium hydroxide	$Ba(OH)_2$	barium hydroxide

On multiple choice section of the AP exam, calculator use is NOT allowed. For many questions, this is a moot point because the question may only involves words, but for other questions, your ability to multiply and factor numbers will be very helpful.

No kidding.....
Practice your multiplication tables.
Go to www.tablestest.com
or www.timestables.me.uk/
or some other multiplication &
division practice site.

This packet should be review. *If you find yourself struggling with any of the topics in the following practices, please email me and let me know where your struggles are, so that I can guide you to more review and practice.*
These topics are review, *and you need to know your stuff so that we can hit the gound running in August.*

This is the periodic Table that you will use in AP Chemistry. Now is as good a time as any to begin to get used to it.

PERIODIC TABLE OF THE ELEMENTS

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.97	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.95	43 **Tc** (97)	44 **Ru** 101.1	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89 **†Ac** (227)	104 **Rf** (267)	105 **Db** (270)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (282)	112 **Cn** (285)	113 **Uut** (285)	114 **Fl** (289)	115 **Uup** (288)	116 **Lv** (293)	117 **Uus** (294)	118 **Uuo** (294)

*Lanthanoid Series	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
†Actinoid Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Summer Review: **Nomenclature Practice** *(Answers on the next page.)*

You should learn your polyatomic ions well enough to write these names and fromulas without looking at the polyatomic ion chart. Use only the periodic table in this packet. (ANSWERS are on the following page)

Write chemical formulas for the following names.

1. sodium sulfite
2. copper(II) nitrate
3. hydrochloric acid
4. sodium hydroxide
5. acetic acid
6. aluminum perchorate
7. silver sulfide
8. carbonic acid
9. ammonium phosphate
10. potassium permanganate
11. lead(II) cyanide
12. calcium acetate
13. nitrous acid
14. hydroiodic acid
15. sodium bicarbonate
16. nickel(III) iodate
17. chloric acid
18. aluminum sulfite
19. phosphorous acid
20. barium hydroxide

Write names for the following formulas.

21. $\mathrm{Sr(CN)_2}$
22. $\mathrm{H_3PO_4}$
23. $\mathrm{ZnSO_4}$
24. $\mathrm{Cu(SO_3)_2}$
25. $\mathrm{H_2SO_4}$
26. AuOH
27. $\mathrm{K_2CO_3}$
28. $\mathrm{NaHCO_3}$
29. $\mathrm{HClO_2}$
30. $\mathrm{AgNO_2}$
31. HBrO
32. KOH
33. $\mathrm{HC_2H_3O_2}$
34. $\mathrm{Ni(BrO_2)_3}$
35. HBr
36. $\mathrm{NaMnO_4}$
37. $\mathrm{HBrO_2}$
38. $\mathrm{H_3PO_4}$
39. $\mathrm{(NH_4)_2SO_4}$
40. $\mathrm{Ni(OH)_2}$

Understanding Net Ionic Equations

Balanced chemical equations, written to represent chemical reactions, are an important part of chemistry.

You do *not* need to write down physical state symbols (aq, s, ppt, L, g ... etc). In fact, it's probably best if you leave them off.

The best way to prepare for writing equations is to practice writing *lots* of equations. Many of the same equation types show up year after year on the AP Exam. When you are reading the words given in a problem, and trying to write an equation, it may be helpful to try to identify the equation as a particular type in order to help you predict the products.

Sometimes you may write overall equations in which all complete chemical formulas are shown. More often, however, equations in AP Chemistry need to be written in **net ionic** form. Net ionic is a term used for balanced equations that describe chemical reactions that occur in aqueous solution. All soluble ionic substances must be written as separated ions with the *spectator ions* left out. The spectator ions are left out of the equation because they do not change form at all during the course of the reaction and do not need to be represented in the chemical equation. All molecular substances and non-soluble compounds must be written as a molecule or formula unit (not ionized!).

Solubility Rules

In first year chemistry we used a solubility chart, however, you will not be allowed to use one in AP Chemistry. The solubility rules that you need to memorize is quite a short list.

ALWAYS SOUBLE IF IN A COMPOUND	EXCEPT WITH
Alkali ions, NH_4^+,	No Exceptions (unless told otherwise in a problem)
NO_3^-, $C_2H_3O_2^-$, ClO_4^-	No Exceptions (unless told otherwise in a problem)
Cl^- , Br^- , I^-	Pb^{2+}, Ag^+ (unless told otherwise in a problem)
SO_4^{2-}	Pb^{2+} (unless told otherwise in a problem)

If a compound does not fit one of the rules above, assume it is INSOLUBLE, unless you are given other information to the contrary within the problem. Non-soluble compounds must be written as a *formula unit* (not ionized). Remember, this list is just a guide, and any information given within a problem that contradicts any rules given above will be followed.

Other considerations to remember when writing chemical equations.

1. Weak acids, (any acid other than the seven strong acids you need to memorize) are mostly *NOT* ionized in solution and thus must be written as molecules. (There are weak bases, you will learn about them during the year.)
2. Strong acids and bases will be considered fully ionized in solution and thus must be written as separated ions in net ionic equations.
3. Soluble salts as memorized from the table above will exist as separated ions in solution and thus must be written as separated ions in net ionic equations, with spectator ions dropping out of the equation.
4. Solids, liquids, and gases should be written as molecules.
5. When you see the words "solution of" to describe an ionic compound, assume that compound is dissolved and dissociated.
6. An ionic compound in a *saturated* solution (saturated: a solution with maximum that can be dissolved) is written in ionic form while an ionic compound in *suspension* (suspension: particles shaken up and floating, but not actually dissolved) should be written together as a molecule or "ionicule."
7. Know your *phantoms* – molecules that when formed as a product, will decompose into a gas and water as indicated below.
   - as a product of a double replacement reaction, H_2CO_3 decomposes into H_2O and CO_2 gas.
   - as a product of a double replacement reaction, NH_4OH decomposes into H_2O and NH_3 gas.

Single Replacement

- A reaction in which one element displaces another in a compound. One element is oxidized and another is reduced. In an oxidation reduction reaction, elements will change their oxidation states.
- Generic: $A + BX \rightarrow B + AX$ or $Y + BX \rightarrow X + BY$
  - **Active metals replace less active metals or hydrogen (in acid or water).**

    The more easily oxidized metal replaces the less easily oxidized metal or hydrogen. You used an activity series in first year chem. You will learn more about that chart, and other methods of predicting which metal is more active than the other.

    - *Magnesium pieces are added to a solution of nickel(III) chlorate.*

      Overall Equation: $3Mg + 2Ni(ClO_4)_3 \rightarrow 2Ni + 3Mg(ClO_4)_2$ $(= 3Mg + 2Ni^{3+} + 6ClO_4^- \rightarrow 2Ni + 3Mg^{2+} + 6ClO_4^-)$

      Net Ionic Equation: $3Mg + 2Ni^{3+} \rightarrow 2Ni + 3Mg^{2+}$

      (Note that the chlorate compounds are soluble and thus the chlorate ions have been removed because they are spectators – unchanged by the reaction. The nickel and magnesium are different as reactants compared to products.)

    - *Nickel is added to hydrochloric acid.*

      Overall Equation: $Ni + 2HCl \rightarrow NiCl_2 + H_2$ $(= Ni + 2H^+ + 2Cl^- \rightarrow Ni^{2+} + 2Cl^- + H_2)$

      Net Ionic Equation: $Ni + 2H^+ \rightarrow Ni^{2+} + H_2$

      (Note that the chloride ions have been removed because nickel(II) chloride is soluble and the chloride ions are in the same form both as reactants and products, thus the chloride ions are spectators – unchanged by the reaction.)

    - *Sodium is added to water.*

      Overall Equation: $2Na + 2H_2O \rightarrow 2NaOH + H_2$

      Net Ionic Equation: $2Na + 2H_2O \rightarrow 2Na^+ + 2OH^- + H_2$

      (Remember that some metals -the alkali metals and some alkaline earth metals- can replace hydrogen in water. You may find it useful to think of water as HOH. For this equation there is no ions that are removed. Na^+ and OH^- must be included as products because those products are not ions on the reactant side.)

  - **Active nonmetals replace less active nonmetals from their compounds in aqueous solution.**

    A halogen will replace a less electronegative (lower on the Periodic table) halogen from their binary salts.

    - *Chlorine gas is bubbled into a solution of potassium iodide.*

      Overall Equation: $Cl_2 + 2KI \rightarrow I_2 + 2KCl$ $(= Cl_2 + 2K^+ + I^- \rightarrow I_2 + 2K^+ + 2Cl^-)$

      Net Ionic Equation: $Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$

      (Note that the potassium ions have been removed because they are spectators – unchanged by the reaction.)

Double Replacement

Two compounds react to form two new compounds. No changes in oxidation numbers occur, thus DR reactions are not redox. Since the movement of electrons does not "push" the reaction, all double replacement reactions must have some other "driving force" that removes a pair of ions from solution. These ions may be removed by forming a precipitate, a gas, or molecular compound. If water forms, the double replacement reaction is an acid/base reaction. If a solid substance forms the double replacement reaction is a precipitation reaction. We can assume that all solutions are aqueous solutions, unless told otherwise.

- **Formation of a precipitate:**
  A precipitate is an insoluble substance formed during the reaction of two aqueous substances. Two ions bond together so strongly that water can not pull them apart. Knowing your solubility rules will help you write these net ionic equations.
  - *Solutions of silver nitrate and sodium sulfide are mixed (Assume a precipitate forms).*

    Overall Equation: $2AgNO_3 + Na_2S \rightarrow Ag_2S + NaNO_3$ $(= 2Ag^+ + NO_3^- + Na^+ + S^{2-} \rightarrow Ag_2S + Na^+ + NO_3^-)$

    Net Ionic Equation: $2Ag^+ + S^{2-} \rightarrow Ag_2S$

    (How do you know which substance is the precipitate? By knowing alkali and nitrate salts are soluble, the precipitate must be the silver sulfide)

- **Formation of a gas:**

  Gases may form from the decomposition of a product such as H_2CO_3 or NH_4OH.

  - *Acetic acid solution is added to a solution of sodium bicarbonate.*

    Overall Equation: $HC_2H_3O_2 + NaHCO_3 \rightarrow NaC_2H_3O_2 + (H_2CO_3 \rightarrow) H_2O + CO_2$

    $(= HC_2H_3O_2 + Na^+ + HCO_3^- \rightarrow Na^+ + C_2H_3O_2^- + H_2O + CO_2)$

    Net Ionic Equation: $HC_2H_3O_2 + HCO_3^- \rightarrow H_2O + CO_2$

    (Note that the acetic acid must be written as a molecule because acetic acid is a weak acid, which you should know from the strong acid chart on page 2. Remember, if an acid is not one of the seven strong acids, you can assume that acid is weak. The carbonic acid, when formed, bubbles off as carbon dioxide. You know this reaction – the classic third grade volcano trick. The sodium ions have been removed because they are spectators – unchanged by the reaction)

  - *A solution of sodium hydroxide is added to a solution of ammonium nitrate.*

    Overall Equation: $NaOH + NH_4NO_3 \rightarrow NaNO_3 + (NH_4OH \rightarrow) NH_3 + H_2O$

    $(= 2Na^+ + OH^- + NH_4^+ + 2NO_3^- \rightarrow Na^+ + 2NO_3^- + NH_3 + H_2O +)$

    Net Ionic Equation: $OH^- + NH_4^+ \rightarrow NH_3 + H_2O$

    (The ammonium hydroxide that is formed, bubbles off as ammonia with water in solution. The sodium and nitrate ions have been removed because they are spectators – unchanged by the reaction)

- **Formation of a molecular substance (often an acid base neutralization):**

  When a molecular substance such as water or a weak acid is formed, ions are removed from solution and the reaction happens. More information on the next page.

Acid base neutralization will be the focus of the next page. → → → → →

Summer Review: Understanding Net Ionic Equations – Acid Base

Acid/Base Neutralization *(a particular "flavor" of double replacement reaction)*

Acids react with bases to produce salts and water.

One mole of hydrogen ions react with one mole of hydroxide ions to produce one mole of water. Remember which acids are strong (and thus ionize completely) and by default, which acids are weak (should be written as a molecule). We can assume that all solutions are aqueous solutions, unless told otherwise.

- *Aqueous solutions of lithium hydroxide and hydrobromic acid are poured together.*

  Overall Equation: $LiOH + HBr \rightarrow H_2O + LiBr$ $(= Li^+ + OH^- + H^+ + Br^- \rightarrow H_2O + Li^+ + Br^-)$

  Net Ionic Equation: $OH^- + H^+ \rightarrow H_2O$

  (A strong acid will be completely ionized in solution – HBr is a strong acid. Lithium bromide is a soluble ionic compound that would be separated into ions. The ions that are unchanged as reactants and products drop out of the equation.)

- *An aqueous solutions of sulfuric acid and barium hydroxide are combined.*

  Overall Equation: $H_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2H_2O$ $(= 2H^+ + SO_4^{2-} + Ba^{2+} + 2OH^- \rightarrow H_2O + Ba^{2+} + SO_4^{2-})$

  Net Ionic Equation: $H^+ + OH^- \rightarrow H_2O$

  (It's true that in the "overall reaction," 2's would show up, but then drop out of the net ionic equation.)

- **Watch out for acids or bases that should be written as a molecule, such as weak acids or weak bases and gases.**

  - *Acetic acid solution is added to a solution of sodium hydroxide.*

    Overall Equation: $HC_2H_3O_2 + NaOH \rightarrow H_2O + NaC_2H_3O_2$ $(= HC_2H_3O_2 + Na^+ + OH^- \rightarrow H_2O + Na^+ + C_2H_3O_2^-)$

    Net Ionic Equation: $HC_2H_3O_2 + OH^- \rightarrow H_2O + C_2H_3O_2^-$

    (Remember, that weak acids are mostly not ionized in solution and thus must be represented as molecules.)

  - *Hydrogen sulfide gas is bubbled through excess potassium hydroxide solution.*

    Overall Equation: $H_2S + 2KOH \rightarrow 2\ H_2O + K_2S$ $(= H_2S + 2OH^- \rightarrow 2\ H_2O + + 2K^+ + S^{2-})$

    Net Ionic Equation: $H_2S + 2OH^- \rightarrow 2\ H_2O + S^{2-}$

    (Remember, that the gas, which is also a weak acid must be written as a molecule.)

  - *A suspension of magnesium hydroxide is added to a dilute solution of hydrochloric acid.*

    Overall Equation: $Mg(OH)_2 + 2HCl \rightarrow 2H_2O + MgCl_2$ $(= Mg(OH)_2 + 2H^+ + Cl^- \rightarrow 2H_2O + Mg^{2+} + 2Cl^-)$

    Net Ionic Equation: $Mg(OH)_2 + 2H^+ \rightarrow 2H_2O + Mg^{2+}$

    (Remember, that a suspension is not actually dissolved, not ionized, and thus must be written as a formula unit – not separated.)

- **Formation of weak acids by combining a weak base with a strong acid.**

  - *Solutions of sodium fluoride and hydrobromic acid are mixed.*

    Overall Equation: $Na^+ + F^- + H^+ + Br^- \rightarrow HF + NaBr$ $(= Na^+ + F^- + H^+ + Br^- \rightarrow HF + Na^+ + Br^-)$

    Net Ionic Equation: $F^- + H^+ \rightarrow HF$

    (Remember that sodium bromide is soluble making the sodium and bromide ions the same as reactants and products, and thus drop out as spectators. The HF is a weak acid, thus must be represented as molecule.)

  - *Solutions of potassium acetate and sulfuric acid are mixed.*

    Overall Equation: $2KC_2H_3O_2 + H_2SO_4 \rightarrow 2HC_2H_3O_2 + K_2SO_4$ $(= 2K^+ + 2C_2H_3O_2^- + 2H^+ + SO_4^{2-} \rightarrow 2HC_2H_3O_2 + 2K^+ + SO_4^{2-})$

    Net Ionic Equation: $C_2H_3O_2^- + H^+ \rightarrow HC_2H_3O_2$

    (Remember $HC_2H_3O_2$ is a weak acid, thus a molecule is formed. Sulfate and potassium ions are spectators. Again, the 2's that show up in the overall and ionic equation will drop out of the net ionic equation.)

Summer Review: Writing Net Ionic Equations – Single Replacement *(Answers on the next page.)* (pg 10 of 26)

Remember that on the AP exam you may only use the periodic table. No solubility chart. Assume that the reaction does occur, thus if you can recognize the single replacement reaction, you do not need to check the activity series. Look for extra information embedded in the question. Answers are on the next page.

1. A strip of magnesium is added to a solution of silver nitrate

2. Aluminum metal is dropped into an solution of zinc chloride

3. Solid silver is dropped into an solution of gold(II) nitrate

4. Aluminum foil is dropped into a solution of nitric acid.

5. Solid barium is added to chlorous acid

6. Potassium metal is dropped into water

7. Liquid bromine is added to an aqueous sodium iodide solution

8. Hydrogen gas is passed over hot copper(II) oxide.

9. Small chunks of solid sodium is added to water.

10. Magnesium metal is added to a dilute solution of nitric acid.

11. Chlorine gas is bubbled into a solution of potassium iodide.

Summer Review: Writing Net Ionic Equations – Double Replacement *(Answers on the next page.)*

Remember that on the AP exam you may only use the periodic table. No solubility chart. Assume that the reaction does occur, thus if you can recognize the double replacement reaction, you should be able to infer the precipitate. Look for extra information embedded in the question. Answers are on the next page.

1. Aqueous solutions of zinc sulfate and sodium phosphate are mixed.

2. Hydrofluoric acid is combined with a solution of lead(II) nitrate.

3. An aqueous solution of lead(II) acetate reacts with hydrochloric acid.

4. Solid sodium carbonate is stirred into hydrobromic acid.

5. Nitric acid is reacted with an aqueous solution of calcium acetate.

6. Hydrochloric acid is poured over powdered potassium carbonate.

7. An aqueous solution of cadmium chloride is reacted with an aqueous solution of potassium phosphate.

8. A solution of hydrofluoric acid is poured over barium carbonate crystals.

9. Hydroiodic acid is poured over potassium carbonate.

10. A solution of sodium hydroxide is poured into a solution of magnesium chloride.

11. Aqueous lead(II) nitrate is combined with potassium iodide.

Summer Review: Writing Net Ionic Equations – Acid Base *(Answers on the next page.)*

Remember that on the AP exam you may only use the periodic table. No solubility chart. Look for extra information embedded in the question. Answers are on the next page.

1. A solution of acetic acid is reacted with a lithium hydroxide solution.

2. A solution of nitric acid is combined with a suspension of magnesium hydroxide.

3. A solution of sulfuric acid is poured over copper(I) hydroxide crystals.

4. A solution of sulfuric acid is added to a solution of barium hydroxide until the same number of moles of each compound has been added, and a precipitate forms.

5. Hydrogen sulfide gas is bubbled through a solution of potassium hydroxide.

6. Potassium hydroxide solution is added to a solution of potassium hydrogen phosphate

7. A solution of sodium hydroxide is added to a solution of sodium dihydrogen phosphate until the same number of moles of each compound has been added.

8. Solutions of sulfuric acid and potassium hydroxide are combined.

9. Hydrochloric acid solution is added to a solution of sodium dihydrogen phosphate

Summer Review: Atomic and Chemical Composition

Navigating all those #'s in the Periodic Table

Atoms	When you look up an element in the periodic chart, and look up its atomic number and mass number, assume you are considering an atom, as opposed to an ion. It is very important to pay close attention to this vocabulary.
Atomic number	tells you the number of protons in an atom. Atoms are neutral in charge, which of course means that the number of protons must equal the number of electrons.
Mass number	is the average atomic mass rounded to the nearest whole number. The mass number is equal to the sum of the protons + neutrons. Thus, to determine the number of neutrons, subtract the atomic number from the mass number.
Ions	During chemical reactions, atoms can lose or gain electrons. In fact they do so on a very regular basis. Since electrons are negatively charged, when electron(s) are lost, an atom turns into an ion and ends up with a positive charge. When electrons are gained, an atom turns into an ion and ends up with a negative charge. Negatively charged ions are called ***anions***. Positively charged ions are called ***cations***.

Symbolizing atoms, isotopes, ions, molecules:

- $_{3}\text{Li}$ the *atomic number* is placed in front of the atom as a subscript.
- ^{7}Li the *mass number* is placed in front of the symbol as a superscript
- Li^{+} the + as a superscript refers to the +1 charge if the atom has turned into an ion
- Li_2 the subscript 2 refers to 2 Lithium atoms that are *stuck* together
- 5 Li the 5 refers to 5 lithium atoms that are NOT stuck together (used as coefficients to balance chemical equations)

Never would all 5 of these numbers be placed around a chemical symbol all at the same time. They would be used at different times in different contexts.

Electron Configuration

Electron configurations are a simple way of writing down the locations of all of the electrons in an atom. Electrons stay within the atom because of their attraction to the protons, they also mutually repel each other, causing them to spread out around the nucleus in regular patterns. This results in geometric areas of probability called **orbitals** (s, p, d, and f) that represent the distinct regions of probability around the nucleus in which each electron exists. The reason that electrons tend to stay in their separate orbitals rather than piling on top of one another is the **Pauli Exclusion Principle**, a theorem from quantum mechanics that dictates that no two electrons can ever be in the same place. The Pauli Exclusion Principle arises from more than just the electrostatic repulsion of negative electrons: it comes from fundamental quantum mechanical principles that constrain all subatomic particles.

The orbitals represent **identifiable "addresses"** for each electron around an atom. Think of the electrons as people going to their favorite concert. The electrons all try to be as close to the stage (the nucleus) as possible, but there is a limited number of seats Some electrons get to be closest to the nucleus, but as the number of electrons that go the concert increases, the further out some of them need to be since the rows closest to the nucleus fill up. This describes a trend observed in the periodic table: elements with small atomic number (and thus fewer electrons) tend to have most of their electrons existing in orbitals near the nucleus. As we move further down the periodic table, orbitals and energy levels further out from the nucleus begin to fill up with electrons. In order to track down where a given electron exists in an atom, you need to know not only how far from the nucleus it is found (described as the electron's **energy level**, since electrons further out from the nucleus tend to have higher energy) but also the *type of orbital* that the electron is be found in.

Example: Arsenic: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$ OR $[\text{Ar}]\ 4s^2\ 3d^{10}\ 4p^3$ (noble gas or condensed version)

You will probably not need to deal with f electrons on the AP exam.

Summer Review: Atomic and Chemical Composition

Percent Composition

Laboratory experiments can give the masses of the various elements contained in the total mass of the compound. This common practice is called elemental analysis or mass percent composition or more simply percent composition. Example, H_2O: 2 (1.01 g/mole) + 16.00 g/mole = 18.02 g/mole total (these numbers are the molar masses from the periodic chart.)

$$H: \frac{2.02g}{18.02} \times 100 = 11.02\% \qquad O: \frac{16.00g}{18.02} \times 100 = 89.88\%$$

Thus water is 11 % hydrogen and 89 % oxygen

Empirical and Molecular Formulas

As you know, the chemical formulas for molecular compounds are not always written in the lowest whole number ratios. We have often used formaldehyde CH_2O and sugar $C_6H_{12}O_6$ as an example. Because of this, the elemental analysis to determine empirical formulas would not allow a chemist to distinguish between sugar and formaldehyde. Another analysis tool, mass spectroscopy would be needed to give one more piece of information: the molar mass of the particular compound being analyzed. So the first 4 steps below will help determine the empirical formula, steps 5-6 must be added to determine the molecular formula.

1. Divide each mass or mass percentage by the molar mass of the element, which will give the number of moles of each element.
2. Divide the results from step 1 by whichever number of moles is the smallest. This maintains the mole ratios from step 1 but bases them on the least abundant element being 1.
3. If some results are far from being whole numbers, multiply all the moles through by a common factor that will convert all the mole amounts to whole numbers or near whole numbers.
4. Round each mole amount to the nearest whole number.
5. If a molar mass is given for the compound, calculate the molar mass for the empirical formula just established from step 4. If the molar mass of the empirical formula is the same as the molar mass of the compound given in the problem, then the empirical formula and the molecular formula are one and the same.
6. If the molar mass of the empirical formula is smaller than the molar mass of the compound, divide the two to determine the whole number factor that the empirical formula must be multiplied by to determine the molecular formula.

Sample Problem

Determine the empirical formula for some compound that was analyzed to be 1.33 g of carbon,0.22 g of hydrogen, and 1.78 g of oxygen. Determine the molecular formula for this compound if the molar mass was measured and found to be 180 g/mole.

- First do step 1 as outlined above.

$$C:\ 1.33g \times \frac{1mol}{12.01g} = 0.111mol \quad H:\ 0.22g \times \frac{1mol}{1.01g} = 0.219mol \quad O:\ 1.78g \times \frac{1mol}{16.00g} = 0.111mol$$

- Proceed to step 2.

$$C:\ \frac{0.111mol}{0.111mol} = 1 \quad H:\ \frac{0.219mol}{0.111mol} = 2 \quad O:\ \frac{0.111mol}{0.111mol} = 1$$

  Voilà. The empirical formula is CH_2O

- Since steps 3 and 4 are not necessary, proceed to step 5
  - For CH_2O molar mass = 30 g/mole which is of course not the same as 180 g/mole
- Proceed to step 6

$$\frac{180}{30} = 6$$

  - Therefore when the factor of 6 is distributed through the empirical formula CH_2O
  - Voilà. The empirical formula converts to $C_6H_{12}O_6$

Summer Review: Atomic and Chemical Composition *(Answers on the next page.)*

1. Complete the following table to demonstrate your knowlege of sub atomic particles

symbol	# of protons	# of neutrons	# of electrons	atomic #	mass #	charge
		24	21			0
			18	15	31	
$^{13}_{6}C$					13	
	17				35	−1
$^{58}_{26}Fe^{3+}$			23		58	

2. Write complete electron configurations for the following particles.
   a. S
   b. Zr
   c. P^{3-}
   d. Cr^{2+}

3. Write condensed electron configurations for the following particles.
   a. Ge
   b. Pb

4. Bismuth subsalicylate, is the active ingredient in Pepto-Bismol which is used to treat upset stomachs. This chemical has the formula $C_7H_5BiO_4$.
   a. Calculate the percent composition of bismuth subsalicylate.
   b. If each tablet of the medication contains 262 milligrams of $C_7H_5BiO_4$ calculate the mass of bismuth in 2 tablets.

5. Determine the empirical and molecular formula of benzene which contains only carbon and hydrogen and is 7.74% hydrogen by mass. The molar mass of benzene is 78.1 g/mol.

6. 6.394 g of compound used as a drying agent is analyzed and determined to be 2.788 g phosphous and 3.606 g oxygen. The molar mass is approximately 284 g/mol. Determine the empirical and moleculoar formulas of this compound. What is the name of this compound.

Summer Review: Stoichiometry

A Typical Plan for Solving Stoichiometry Problems

There is a basic pattern to all stoichiometry problems, with variations depending on what information is given and what questions must be answered. You are using dimensional analysis so be sure to set up you calculations with the starting units on top and bottom so it will cancel out and with the desired substance on the top.

A. You must start with a balanced equation.
B. Convert the units of any starting substances into moles. (USE Molar Mass (g/mol) complete this calculation) Since the stoichiometric LINK or RATIO – coefficients from the balanced equation – is in moles, you must work the problem in moles.
C. Reread the problem to determine the information that you need to calculate. Use the stoichiometric LINK to convert from a known substance to a desired substance that you need to answer the question. Note that the LINK is set up with the known substance on the bottom (so it will cancel out) and with the desired substance on the top.
D. If necessary, convert any answers back into grams.

If your problems only involve only moles, then you can skip steps B and E

Sample Problem

Lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can be absorbed by 1.00. kg of lithium hydroxide?

STEP A. balanced equation: $2\ LiOH + CO_2 \rightarrow Li_2CO_3 + H_2O$

Notice that the starting info is given in kilograms, so 1.00 kg should be converted to grams.

$$1.00\times10^3 gLiOH\times\frac{1mol}{23.99g}\times\frac{1CO_2}{2LiOH}\times\frac{44.01gCO_2}{1molCO_2}=917gCO_2$$

Steps in the dimensional analysis: B. C. D.

STEP B. Change to moles using the molar mass of LiOH
STEP C. Change from moles of LiOH to moles of CO_2 using the coefficients from the balanced equation.
STEP D. Change from moles of CO_2 back to grams of CO_2 using the molar mass of CO_2.

Problem Solving Plan - Limiting Reactant and Percent Yield

For limiting reactant problems, the problem will give you information about two reactants as opposed to information given for only reactiant and an assumption that the other reactant is present in excess.

A. You must always start with a balanced equation.
B. If it is a limiting reactant problem.... Determine which reactant LIMITS
C. First you *must* change your mass values to moles. NOTE: The mathematical trick to determine which reactant limits is to divide the moles of each reactant by the coefficient (from the balanced equation) associated with that reactant. The number that comes out the smallest indicates which reactant is the limiting one. The limiting reactant is the one that you must base all your other calculations on because it is the substance that limits how much of everything else can be made or is needed.
D. Solve the problem using the same steps for Stoichiometry problems above based on the LR.
E. Of course, the other reactant (if there's only two) will be the excess reactant, and some of it will be left over. (Knowing which reactant limits and which is excess, use the limiting reactant to set up a stoichiometric LINK to determine the mass of the excess reactant that is actually needed to do the reaction. Then, subtract the mass of reactant that you just calculated was needed from the amount of excess reactant started with to determine the mass of excess reactant that is left over.
F. Determining Percent Yield- After determining the LR, use the link to calculate the theoretical amount of the product for which you need a yield.
G. The experimental amount actually produced will be given in the problem. Use it to set up the equation below and determine the percent yield:

$$\frac{Experimental}{Theoretical}\times100=PercentYield$$

Molarity (M):

This is the most common method of reporting concentration used in AP chemistry.

Molarity is the number of moles of solute per liter of solution. $Molarity = \frac{MolesofSolute}{LitersofSolution}$

When a more concentrated solution is diluted, the moles of the solute will be the same before and after the dilution. This gives rise to the dilution equation, which is just a variation of the molarity equation. $M_cV_c = M_dV_d$

Sample Problems

1. What is the molarity of a solution that contains 6.57 g of magnesium chloride in 250. mL of solution?
    - *First you need to be able to write out the chemical formula for maganesium chloride, and calculate the molar mass.*
        - $MgCl_2$ 24.31 + 2 × 35.45 = 95.21 g/mol
    - *Next, convert the mass of magnesium chloride to moles.* $6.57 gMgCl_2 \times \frac{1 molMgCl_2}{95.21 gMgCl_2} = 0.0690 molMgCl_2$
    - *Next apply the molarity equation.* $\frac{0.069 mol}{0.25 L} = 0.276 M$

Hey look! Molarity which is moles per liter, is also millimoles per milliliter!!

$$\frac{5 mol}{1L} \times \frac{1000 millimol}{1 mol} \times \frac{1L}{1000 mL} = \frac{5 millimol}{1 mL}$$

2. Given 25.0 mL of a 0.05 M of aluminum sulfate solution.
    (a) How many millimoles of aluminum sulfate does this solution contain?
    (b) How many millimoles of sulfate does this solution contain?
    - *To answer (a), simply apply the molarity equation* $M \times V = moles$ $0.05 M \times 25 ml = 1.25 millimol$
    - *To answer part (b) you need to write out the chemical formula for aluminum sulfate.* $Al_2(SO_4)_3$
    - *Thus you can see there are three sulfate ions per aluminum sulfate.*
        - $1.25 mmolAl_2(SO_4)_3 \times \frac{3SO_4^{2-} ions}{1Al_2(SO_4)_3} = 3.75 mmol\ SO_4^{2-} ions$

3. If 38.0 mL of a 6.0 M HCl solution are diluted to a final volume of 250 mL, what is the final concentration?
    - *To answer (a), simply apply the dilution equation* $M_cV_c = M_dV_d$
        - $M_cV_c = M_dV_d$ $M_d = \frac{M_cV_c}{V_d}$ $M_d = \frac{6M \times 38 mL}{250 ml}$ $M_d = 0.91M$

Summary Review: Stoichiometry *(Answers on the next page.)*

1. Solutions of nickel(II) chloride and potassium phosphate will react to produce a light green precipitate .
   a) Write a balanced overall chemical equation to represent this reaction.
   b) What mass of potassium phosphate in solution would be required to react completely with 0.875 g of nickel(II) chloride in solution?
   c) Calculate the theoretical mass of nickel(II) phosphate that could be produced.
   d) Convert the overall equation to the net ionic equation.

Solution	Concentration (Molarity)	Volume (Liters)
$Na_2S_2O_3$	0.500	250.
NOCl	2.00	150.
NaOH	0.600	175

2. Gallium metal reacts with perchloric acid. *Assume that at room conditions, 24.0 L is the volume of 1.00 mole of gas.*
   a) Write both overall and net ionic balanced equations to represent this reaction.
   b) If 2.25 L of hydrogen gas were collected, what mass of gallium metal was dropped into the acid solution?

3. Aluminum will cause copper to reduce from a solution of copper(II) chloride.
   a) Write a balanced net ionic chemical equation to represent this reaction.
   b) Is 5.00 g of aluminum enough aluminum to reduce all of the copper(II) ions from 750. ml of a 0.500 M solution?
   c) If 5.00 g of aluminum is more than enough, what mass would be left over? OR if 5.00 g of aluminum is not enough, what is the additional mass of aluminum that would be needed to remove all of the copper(II) ions from solution?

4. Hydrochloric acid reacts with solid magnesium hydroxide.
   a) Write a balanced overall chemical equation to represent this reaction.
   b) What volume, in milliliters, of 0.25 M hydrochloric acid solution would be required to completely react with 4.56 g of magnesium hydroxide?
   c) Convert the balanced overall equation to a net ionic equation.

5. 1.65 g of zinc is dropped into 150. ml of 0.250 M of hydrobromic acid.
   a) Write both overall and net ionic balanced chemical equations to represent this reaction.
   b) Which reactant is the limiting reactant in this chemical reaction?
   c) Calculate the theoretical mass of solid zinc bromide that should be produced.
   d) If Consuela and Pete were able to produce 3.67 g of the zinc bromide, what is their percent yield?

6. Eldon and Sally were preparing a sulfuric acid solution for a lab and they needed 500. ml of 0.045 M
   a) Calculate the volume of 3.0 M solution that Eldon and Sally should measure out into the 500. ml volumetric flask.
   b) What is the molarity of H^+ ions for the solution that Eldon and Sally prepared?
   c) What are the number of millimol of H^+ ions that are in Eldon and Sally's 500. ml of 0.045 M sulfuric acid solution?

7. Nitric acid will react with a sodium carbonate solution.
   a) Write a balanced overall equation to represent this reaction. (Hint: one of the products is a phantom, and will turn into two products. Refer to page 5 of this packet for more information.)
   b) What volume of 0.25 M nitric acid would be required to react completely with 245 ml of 0.38 M of the sodium carbonate solution.

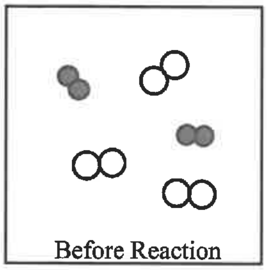
$$Na_2S_2O_{3(aq)} + 4\ NaOCl_{(aq)} + 2\ NaOH_{(aq)} \rightarrow 2Na_2SO_{4(aq)} + 4\ NaCl_{(aq)} + H_2O_{(L)}$$

8. Answer the following questions about the balanced redox equation shown above.
   a) The student combines the solutions shown in the table to the right. Determine the limiting reactant.
   b) How many moles of water would be produced during this reaction?
   c) Convert the overall equation shown above into a net ionic equation.

Summary Review: Particulate Diagrams *(Answers on the next page.)*

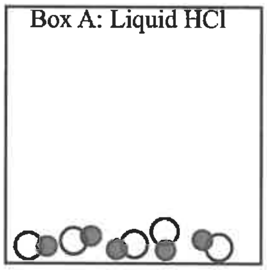
The AP Chemistry curriculum has made a point of asking students to interpret and draw particulate diagrams. A particulate diagram is a sketch that asks students conceptualize what may be happening at the atom and molecule level. You will be asked to convert between macroscopic observations in lab, to symbolic representations with chemical formulas and balanced equations, to particulate representations of the atoms, ions and molecules.

1. The picture shown to the right is a representation of a mixture of hydrogen and oxygen molecules that can be sparked to produce water. Draw a sketch that represents the resulting mixture after the reaction goes to completetion.
   Hint: write a balanced chemical equation first. Decide which molecule best represents oxygen and which best represents hydrogen.



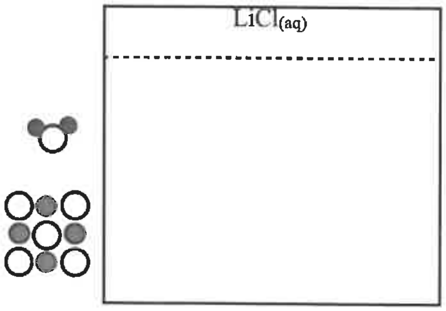
After Reaction

2. Draw a sketch that represents five molecules of HCl in the liquid state is shown in Box A below. In Box B, draw a representation of the five molecules of HCl after complete vaporization has occurred



Box B: After Complete Vaporization

3. A section of a solid lithium chloride crystal is represented to the left of the box below. In the box, show the interactions of the components of a lithium chloride crystal dissolved in water by making a drawing that represents the different particles present in the solution. Include only **one formula unit** of lithium chloride and at least three, but no more than five molecules of water. Your drawing must include the following details.
   - identify the ions (symbol and charge)
   - the proper arrangement and orientation of the particles in the solution



$N_2 + 3F_2 \rightarrow 2NF_3$

4. The picture shown to the right is a representation of a mixture of ammonia and hydrogen molecules that is a result of the completion of the reaction between N_2 and H_2 as shown in the reaction above. In the box on the left, draw the particle-level representation of the reactant mixture of N_2 and H_2 that would yield the product mixture shown in the box on the right. In your drawing, represent nitrogen atoms and hydrogen atoms as shown below

● = fluorine
○ = nitrogen

Before Reaction