

HANDOUT SET

GENERAL CHEMISTRY I

Periodic Table of the Elements

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	IA												IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1 H 1.00794																	2 He 4.00262
2	3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.179
3	11 Na 22.9898	12 Mg 24.305											13 Al 26.98154	14 Si 28.0855	15 P 30.97376	16 S 32.066	17 Cl 35.453	18 Ar 39.948
4	19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.59	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
5	37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.41	49 In 114.82	50 Sn 118.710	51 Sb 121.75	52 Te 127.60	53 I 126.9045	54 Xe 131.29
6	55 Cs 132.9054	56 Ba 137.34	57 La* 138.91	72 Hf 178.49	73 Ta 180.9479	74 W 183.85	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.9665	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.9804	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra 226.0254	89 Ac** 227.0278	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (264)	108 Hs (265)	109 Mt (266)	110 (270)	111 (272)	112 *** (277)						

*Lanthanides	58 Ce 140.12	59 Pr 140.9077	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.9342	70 Yb 173.04	71 Lu 174.967
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**Actinides	90 Th 232.038	91 Pa 231.0659	92 U 238.0289	93 Np 237.0482	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 (260)
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Mass numbers in parenthesis are the mass numbers of the most stable isotopes. As of 1997 elements 110-112 have not been named.

***Peter Armbruster and Sigurd Hofman synthesized a single atom at the Heavy-Ion Research Center in Darmstadt, Germany in 1996. The atom survived for 280 μ s after which it decayed to element 110 by loss of an α -particle

Chapters 4 and 5

Chemical Reactions
and
Reaction Stoichiometry

CHEMICAL REACTIONS: REACTION STOICHIOMETRY

CHAPTER 4

INTRODUCTION Chemical reactions are common experiences for us yet we rarely think about the process itself or that quantities of reactants and products are involved. This chapter starts by introducing the *chemical equation*, a shorthand way of describing a chemical reaction including the mass (or mole) balance. Nomenclature plays an important role in writing chemical equations so earlier studies of chemical formulas and nomenclature tie directly into this aspect of chemistry. After writing a correct and balanced chemical equation, it is then possible for the scientist (that would be you) to perform calculations involving masses or moles of reactants and products. Thus, it is possible to calculate the mass of the product expected from a given amount reactant and *vice versa*. Often one reactant is in limiting quantity so some amount of it remains after the reaction is complete. Through stoichiometric calculations, it is possible to calculate the amount of material remaining in a reaction involving limiting reagents. Finally, since many reactions are performed using solutions, the concept of concentration (often in molarity) is introduced and used.

- GOALS**
1. You must be able to complete and balance a variety of chemical reaction types. It is especially important that you can write and balance a reaction from the names of the reactants and products. Remember, a complete reaction also has the states of matter associated with each reactant and product.
 2. Mass-to-mass (mole-to-mole) calculations are at the heart of the sciences which use chemistry as their foundation. It is essential that you can handle virtually any stoichiometric calculation. Remember, stoichiometry revolves around the mole ratio and the conversion of the moles of known substance to the moles of the substance desired.
 3. Stoichiometric calculations are no different for problems involving one reactant in limiting quantity but it is important that you are able to recognize and do calculations involving limiting reagents.
 4. Molarity as a concentration unit is important in performing stoichiometric calculations in situations such as simple solution chemistry and titration. More concentrated solutions may be diluted. You should know how to perform this operation and determine the new concentration of the solution after dilution.
 5. Sometimes not all of the expected product in a reaction is realized experimentally. The percentage yield is a measure of this difference.
 6. Nomenclature is becoming increasingly more important, especially in writing reactions.

DEFINITIONS

You should have a working knowledge of at least these terms and any others used in lecture.

Chemical equation
Chemical reaction
Chemical balance
Limiting reagent
Titration

Mole ratio
Concentration
Molarity
Molar concentration
Dilution

Solute
Solvent
Solution
Theoretical yield
Actual yield
Percentage yield

REACTIONS IN AQUEOUS SOLUTION

CHAPTER 5

INTRODUCTION We have been introduced to a wide array of chemical reactions and reaction types. Additionally, we have seen that substances can fall into 3 different classes of electrolytes; weak, strong, and non-electrolytes. These concepts together and others as well allow us to predict reaction products and determine the reagents necessary to form a desired product. Of particular importance is writing and balancing oxidation-reduction reactions. The instructor recommends the balancing technique outlined in lecture over those of the authors of the textbook, but the textbook method works equally well. Finally, while titration was discussed in the prior chapter and laboratory, the formal introduction came in chapter 5.

- GOALS**
1. You should be able to write and balance chemical reactions of *all* reaction types presented in lecture and lab from formulas *and* names. Nomenclature is especially important in this chapter.
 2. You should be able to identify the species by type (oxidized, oxidizing agent, reduced, etc.) in oxidation-reduction reactions.
 3. The concept that compounds can be weak and strong electrolytes or non-electrolytes is critical in writing and predicting reactions.
 4. Knowledge of the activity series and its use is important for specific reaction types - single-displacement.
 5. Understanding titration and the stoichiometric calculations involved is very important.

DEFINITIONS

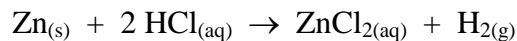
You should have a working knowledge of at least these terms and any others used in lecture.

Acid	Single-displacement	Whole ionic equation
Strong acid	Double-displacement	Net ionic equation
Weak acid	Metathesis	Precipitate
Base	Combustion	Titration
Strong base	Neutralization	Titrant
Weak base	Oxidation-reduction reaction	Burette
Salt	Oxidation	Volumetric flask
Electrolyte	Reduction	Pipette
Strong electrolyte	Oxidizing agent	Indicator
Weak electrolyte	Reducing agent	Equivalence point
Nonelectrolyte	Whole molecular equation	Endpoint
Combination reaction		Standardization
Decomposition reaction		

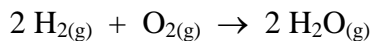
Reaction Stoichiometry I

Mass-Mole Relationships and Limiting Reactant

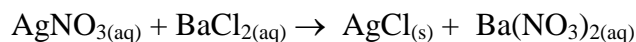
1. According to the equation given, what mass of H₂ can be produced by the reaction of 10.0 g of Zn with the hydrochloric acid?



2. According to the equation of the reaction of H₂ with O₂, what mass of water can be made from the mass of hydrogen from question 1?



3. For the equation...



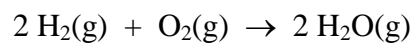
a) Balance the equation

b) How many moles of silver chloride can be produced starting with 2.0 moles of AgNO₃? ...1.5 moles BaCl₂?

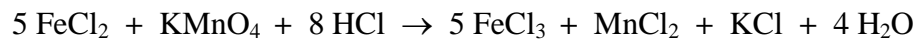
c) How many mol of AgCl will be produced if 3.00 g of BaCl₂ react with a sufficient amount of AgNO₃?

d. What mass of BaCl_2 is necessary to form 5.15 g of AgCl ?

4. What mass of water would be formed from the reaction of 5.0 g H_2 , and 30.0 g O_2 ? How much of the excess reagent remains unreacted?



5. Calculate the mass of iron(III) chloride produced when 1.05 g iron(II) chloride reacts with 1.10 g of potassium permanganate (KMnO_4) according to the equation...



Reaction Stoichiometry II

Extra Stoichiometry Study Questions

- What quantity, in moles, of water are formed by the reaction of 0.112 mol of oxygen gas with excess hydrogen gas?
 - In the same reaction as above, what quantity in moles of hydrogen gas are consumed? (Just for fun... how many molecules of hydrogen gas were consumed?)
 - If 1.000 mol of hydrogen gas was available, how much remains?
- The reaction of silver nitrate with potassium chloride yields silver chloride as one of the products. What mass, in grams, of silver chloride will be produced by the reaction of 1.00 g of silver nitrate with 0.439 g of potassium chloride?

3. Assume that gasoline is entirely C_7H_8 (toluene). What mass of CO_2 will be produced by burning 1 tank (48.0 L) of gasoline in your car? The density of toluene is 0.866 g/mL. (FYI, gasoline is an enormously complex mixture of aliphatic and aromatic hydrocarbons with the toluene being a major component.)

Reaction Stoichiometry Parallel Reactions

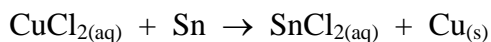
Petrucci & Harwood, Chapter 4

The manufacture of ethyl alcohol, $\text{C}_2\text{H}_5\text{OH}$, yields diethyl ether, $(\text{C}_2\text{H}_5)_2\text{O}$, as a by-product. The complete combustion of a 1.005-g sample of the product of this process yields 1.963 g CO_2 . What must be the mass percents of $\text{C}_2\text{H}_5\text{OH}$ and of $(\text{C}_2\text{H}_5)_2\text{O}$ in this sample?

Reaction Stoichiometry III

Yield, Concentration, and Solution Stoichiometry

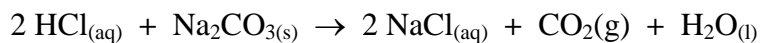
1. If 5.05 g of CuCl_2 reacts with excess tin, what mass of copper is expected from the reaction?



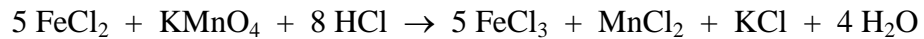
What is the percentage yield if only 2.20 g of Cu is recovered? Where did the lost product go?

2. What is the molar concentration of a solution prepared by dissolving 5.84 g of NaCl in sufficient water to make 1.0 L of solution?
3. What mass of NaOH is necessary to prepare 500.0 mL of 0.250 M NaOH? (molar mass of NaOH = 40.00 g/mol)

4. What quantity, in moles, of HCl are present in 10.00 mL of 0.125 M HCl? What would the new HCl concentration be if 10.00 mL of water were added to the original 10.00 mL of 0.125 M HCl?
5. What is the concentration of the solution prepared by diluting 10.00 mL of 0.125 M HCl to 40.00 mL?
6. How might you prepare 500.0 mL of 0.10 M HCl from concentrated HCl (12 M)?
7. What volume (in mL) of 0.15 M HCl will react with 2.50 g of Na₂CO₃ according to the equation



8. What is the concentration of a KMnO_4 solution if 15.16 mL of the KMnO_4 solution is required to titrate to equivalence point 25.00 mL of 0.08765 M FeCl_2 ?



9. What is the percentage purity of a sample of impure oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$, 90.04 g/mol) if 23.42 mL of the KMnO_4 from problem 8 was required to titrate 1.5960 g of the solid?

Concentration Calculation Involving Parts-per-Million

A solution used to chlorinate a home swimming pool contains 7% chlorine by mass. An ideal chlorine level for the pool is one part per million (1 ppm). If you assume densities of 1.10 g/mL for the chlorine solution and 1.00g/mL for the swimming pool water, what volume of the chlorine solution, in liters, is required to produce a chlorine level of 1 ppm in a 18,000-gallon swimming pool?

Oxidation Number Rules

1. The oxidation number of the element in its most stable form is zero.
2. The oxidation number of a monatomic ion is its ionic charge.
3. The sum of the oxidation numbers in a...
molecule is zero
polyatomic ion is the charge of the ion
4. Except in hydrides, hydrogen in a compound or polyatomic ion is +1
5. Except in peroxides and superoxides, oxygen in a compound or polyatomic ion is -2

Peroxide: O_2^{2-} each O = -1

Superoxide: O_2^- each O = $-\frac{1}{2}$

The Activity Series for Metals

Li			
K			
Ba			
Ca			
Na			
Mg			
Al			
Zn			
Cr			
Fe			
Cd			
Co			
Ni			
Sn			
Pb			
H			
Cu			
Hg			
Ag			
Pt			
Au			

General rule-of-thumb: The neutral form of the metal (or hydrogen) will react with the ion form of the metal (or hydrogen ion) below itself.

BALANCING OXIDATION-REDUCTION REACTIONS

There are two methods to balance redox reactions: 1) **Oxidation Number Method** and 2) **Half-reaction (Ion-electron) Method**. Both have their strengths and weaknesses. The oxidation number method is very fast while the half-reaction method allows the use of tables of electrochemical half-cell potentials to be used to determine whether a reaction will proceed. In this problem session, an algorithm will be given for balancing by the oxidation number method.

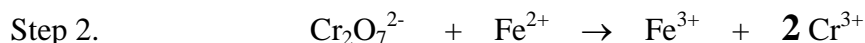
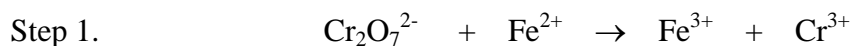
OXIDATION-NUMBER METHOD

This method revolves around assigning oxidation numbers to the species whose oxidation number changes during the redox reaction.

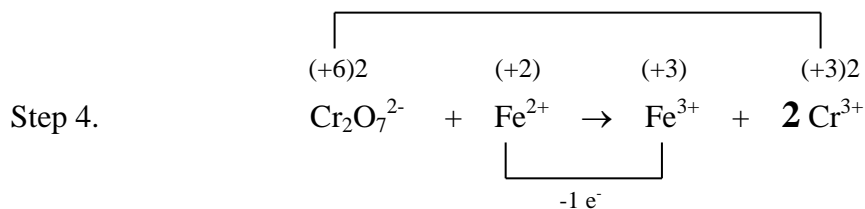
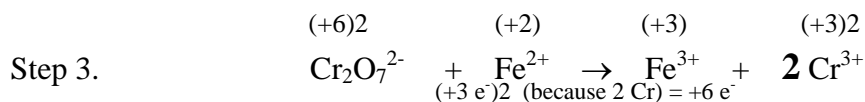
1. Write the reaction from words if necessary. Ignore spectator ions.
2. If necessary, adjust the stoichiometric ratios of the species. This is important so that the final species balance is correct. An illustration of this “pre-balance” is shown in the example problem.
3. Assign oxidation numbers to atoms which undergo a change in oxidation number. Connect identical atoms across the arrow with a line and place the number of electrons gained or lost by the species on the line.
4. Equalize the electrons to the least-common-multiple with multipliers. Multiply the stoichiometric coefficients for the species connected by the lines by the multipliers.
5. Balance the ionic charges on the left and right side of the arrow with either H^+ or OH^- such that the sum of all the charges on the left and right are equal. Unless otherwise specified, usually (but not always), hydrogen ion is used when transition metals are involved and hydroxide is used when nonmetals are the only species in the reaction. If there are no ions in the reaction, this step is skipped.
6. Balance the addition of H^+ or OH^- with H_2O as needed.
7. Check the overall balance. If everything is correct, place the spectator ions back into the equation as necessary.

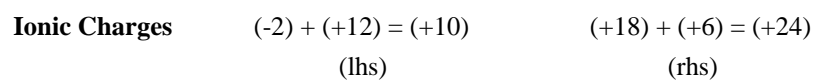
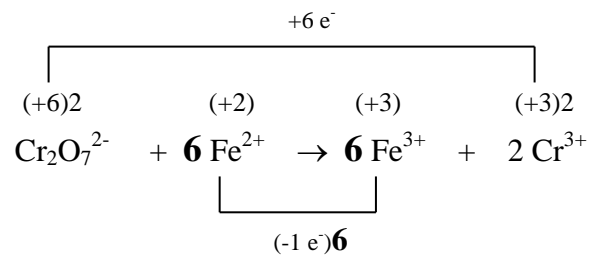
Example

Dichromate ion reacts with iron(II) ion to form chromium(III) ion and iron(III) ion. Write and balance the reaction.



The 2 needs to be added here so that the ratio of chromium on the left and right are correct.

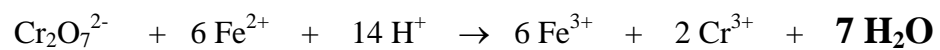




Step 5. Since transition metals involved, charge balance with acid...



Step 6. Finally, balance with water on the right...



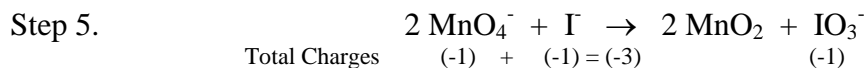
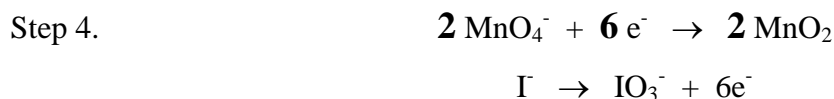
HALF-REACTION METHOD

This method splits the oxidation and reduction reactions apart and balances them separately. The balanced reactions are then added together to give the overall reaction.

1. Write the reaction from words if necessary. Ignore spectator ions.
2. Determine by oxidation number which species are oxidized and which are reduced.
3. Write half-reactions showing the oxidation reaction and the reduction reaction. Be sure that the redox species are balanced; Do not worry about other atoms (such as oxygen, unless it is one of the redox species).
4. Balance the electrons so that each reaction gains or loses the same number of electrons. Multiply the redox species stoichiometric coefficient by the multipliers used to equalize the electrons.
5. Add the two half-reactions together. Since the number of electrons on the left and right side of the arrow are the same they may be cancelled.
6. Balance the charge with hydrogen ion or hydroxide ion. Unless otherwise specified, usually (but not always), hydrogen ion is used when transition metals are involved and hydroxide is used when nonmetals are the only species in the reaction. If there are no ions in the reaction, this step is skipped.
7. Balance the mass with water as in the oxidation-number method
8. Check the overall balance and replace the spectator ions as necessary.

Example

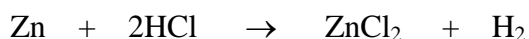
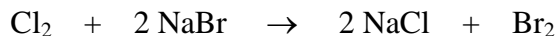
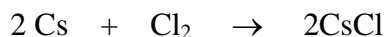
Complete and balance the reaction of potassium permanganate with potassium iodide in slightly basic solution. which forms manganese(IV) oxide and iodate ion.



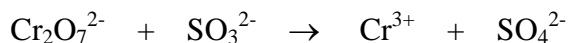
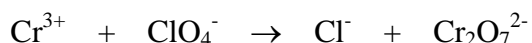
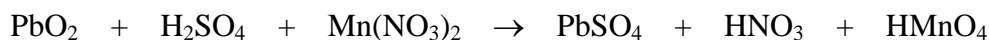
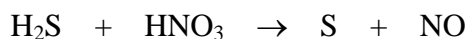
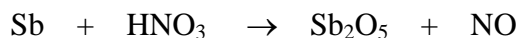
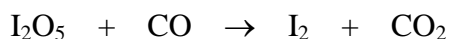
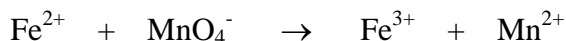
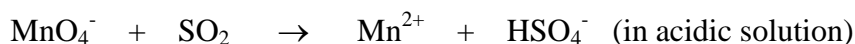
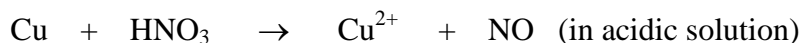
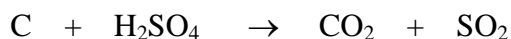
Additional RedOx Reactions

Refer to *Balancing Oxidation-Reactions* while working on these problems. You should also refer to your textbook chapter on oxidation-reduction reactions.

1. For the complete redox reactions below (a) break down each reaction into a set of half-reactions and (b) identify the oxidizing and reducing agents.



2. Balance the following redox reactions. Work the problems on a separate sheet of paper.



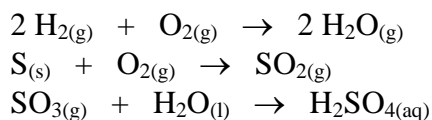
Extra Stoichiometry Study Questions

1. What quantity, in moles, of water is formed by the reaction of 0.112 mol of oxygen gas with excess hydrogen gas?
2. In the same reaction as above, what quantity in moles of hydrogen gas are consumed? (Just for fun... how many molecules of hydrogen gas were consumed?)
3. If 1.000 mol of hydrogen gas was available, how much remains?
4. The reaction of silver nitrate with potassium chloride yields silver chloride as one of the products. What mass, in grams, of silver chloride will be produced by the reaction of 1.00 g of silver nitrate with 0.439 g of potassium chloride?

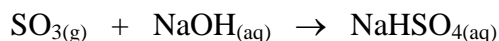
5. Assume that gasoline is entirely C_7H_8 (toluene). What mass of CO_2 will be produced by burning 1 tank (48.0 L) of gasoline in your car? The density of toluene is 0.866 g/mL. (FYI, gasoline is an enormously complex mixture of aliphatic and aromatic hydrocarbons with the toluene being a major component.)

Types of Chemical Reactions

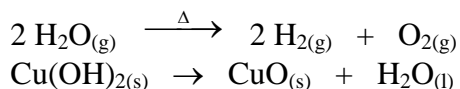
Combination



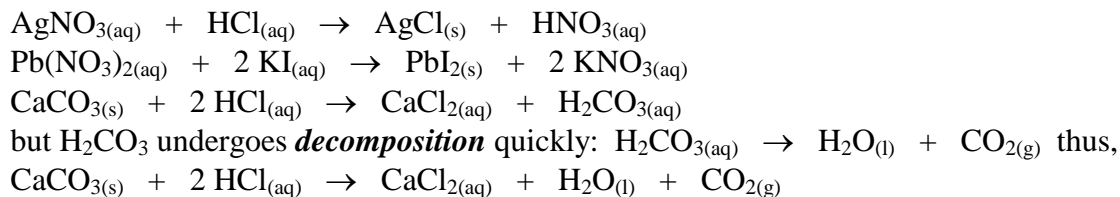
A special case of combination – Acid anhydride-Base Neutralization



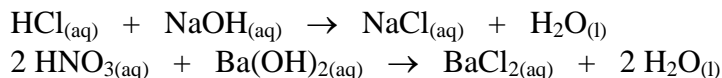
Decomposition



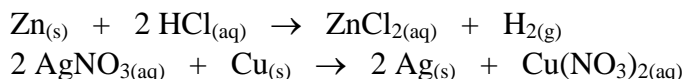
Metathesis (Double-Displacement)



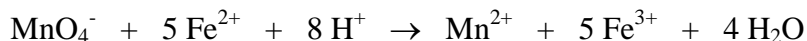
A special case of metathesis - Acid-Base Neutralization



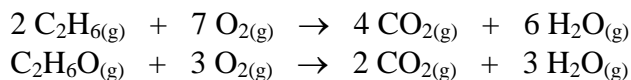
Single-Displacement



Oxidation-Reduction



Combustion



SOLUBILITY RULES FOR IONIC COMPOUNDS

The following table outlines the general rules for predicting the solubility and electrolyte nature of ionic compounds in aqueous solution. These rules are also helpful in predicting the outcome of some reactions of ionic compounds in aqueous solution.

In general, a chemical reaction involving ionic substances will occur if a precipitate, gas, or weak electrolyte is formed as a product. The rules given are for aqueous solution and for ionic compounds only. The rules are not exhaustive and exceptions exist. The rules are not applicable for nonaqueous solutions or molecular compounds.

In general...	
1. Group IA compounds including NH_4^+	All compounds are soluble and strong electrolytes.
2. Compounds of NO_3^- , CH_3CO_2^- , ClO_4^-	All compounds are soluble and strong electrolytes.
3. Compounds of Cl^- , Br^- , I^- , CN^- , SCN^-	All compounds are soluble and strong electrolytes except those of Ag^+ , Hg_2^{2+} , and Pb^{2+} .
4. Compounds of SO_4^{2-}	All compounds are soluble and strong electrolytes except those of Pb^{2+} , Hg_2^{2+} , and Group IIA (Ca to Ra). CaSO_4 and Ag_2SO_4 are sparingly soluble.
5. Compounds of OH^- and S^{2-}	All compounds are insoluble and weak or nonelectrolytes except those of Group IA and Group IIA (Ca to Ra). Group IIA sulfides, Al_2S_3 , and Cr_2S_3 all decompose in aqueous solution.
6. Compounds of CO_3^{2-} , PO_4^{3-} , AsO_4^{3-} , CrO_4^{2-}	All compounds are insoluble and weak or nonelectrolytes except those of Group IA and NH_4^+ .

The formation of carbonic acid in a reaction usually results in the spontaneous decomposition:



The formation of sulfurous acid (not sulfuric acid) also produces an analogous reaction:



For this reason, CO_2 and SO_2 are two examples of acidic anhydrides.