KEY

GENERAL & ANALYTICAL CHEMISTRY I
CHMG-141
With Dr. Bailey

Name___________

Recitation
Week11
Chemical Reactions

SOLUBLE COMPOUNDS

	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^- chlorate, ClO_3^- perchlorate, ClO_4^- acetate, $CH_3CO_2^-$	
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Salts of sulfate, SO_4^{2-}	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

INSOLUBLE COMPOUNDS

	EXCEPTIONS
Most salts of carbonate, CO_3^{2-} phosphate, PO_4^{3-} oxalate, $C_2O_4^{2-}$ chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	Ba $(OH)_2$ is soluble

Rules for Assigning Oxidation States

In their compounds, nonmetals have oxidation states according to the table below

✓ nonmetals higher on the table take priority

Nonmetal	Oxidation State	Example
F	-1	CF_4
H	+1	CH_4
O	-2	CO_2
Group 7A	-1	CCl_4
Group 6A	-2	CS_2
Group 5A	-3	NH_3

21

Problem 1:

2) Complete and balance each of the following equations. If no reaction occurs, write NO REACTION:

a. $NaNO_3(aq) + KCl(aq) \rightarrow$ **NO REACTION** *(all reactants and possible products are soluble)*

$NaNO_3(aq) + KCl(aq) \rightarrow NaCl(aq) + KNO_3(aq)$

Complete ionic equation

$Na^+(aq) + NO_3^-(aq) + K^+(aq) + Cl^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + K^+(aq) + NO_3^-(aq)$

All the ions are spectator ions

b. $2K_3PO_4(aq) + 3NiCl_2(aq) \rightarrow 6KCl(aq) + Ni_3(PO_4)_2(s)$

REACTION OF PRECIPITATON

Complete ionic equation:

$6K^+(aq) + 2PO_4^{-3}(aq) + 3Ni^{+2}(aq) + \cancel{6Cl^-(aq)} \rightarrow \cancel{6K^+(aq)} + \cancel{6Cl^-(aq)} + Ni_3(PO_4)_2(s)$

Net ionic equation:

$$\boxed{3Ni^{+2}(aq) + 2PO_4^{-3}(aq) \rightarrow Ni_3(PO_4)_2(s)}$$

c. $NH_4Cl(aq) + AgNO_3(aq) \rightarrow NH_4NO_3(aq) + AgCl(s)$

REACTION OF PRECIPITATON

Complete ionic equation:

$\boxed{Cl^-(aq)} + NH_4^+(aq) + \boxed{Ag^+(aq)} + NO_3^-(aq) \rightarrow \cancel{NH_4^+(aq)} + \cancel{NO_3^-(aq)} + \boxed{AgCl(s)}$

Net ionic equation:

$$\boxed{Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)}$$

d. $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$

REACTION OF NEUTROLIZATION

Complete ionic equation:

$Na^+(aq) + \boxed{OH^-(aq)} + \boxed{H^+(aq)} + \cancel{NO_3^-(aq)} \rightarrow \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)} + \boxed{H_2O(l)}$

Net ionic equation:

$$\boxed{OH^-(aq) + H^+(aq) \rightarrow H_2O(l)}$$

Gas forming reaction e) $K_2S(aq) + 2HCl(aq) \rightarrow H_2S(g) + 2KCl(aq)$

Problem 2:

a) Write Balanced complete ionic and net ionic equation for each of the following reactions:

a. $K_2SO_4(aq) + CaCl_2(aq) \rightarrow CaSO_4(s) + 2KCl(aq)$ – **Molecular Equation; REACTION OF PRECIPITATON**

$\cancel{2K^+(aq)} + SO_4^{2-}(aq) + Ca^{2+}(aq) + \cancel{2Cl^-(aq)} \rightarrow CaSO_4(s) + \cancel{2K^+(aq)} + \cancel{2Cl^-(aq)}$ – **Complete Ionic Equation**

$\boxed{Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)}$ – **Net Ionic Equation**

b. $HCl(aq) + LiOH(aq) \rightarrow H_2O(l) + LiCl(aq)$ – **Molecular Equation; REACTION of NEUTROLIZATION**

$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Li^+(aq)} + OH^-(aq) \rightarrow \cancel{Li^+(aq)} + \cancel{Cl^-(aq)} + H_2O(l)$ — **Complete Ionic Equation**

$\boxed{H^+(aq) + OH^-(aq) \rightarrow H_2O(l)}$ - **Net Ionic Equation**

~~b) HCl(aq) + LiOH(aq) → H_2O(l) + LiCl(aq)~~

Part B

Problem 1: Determine the oxidation number for each element in the following compounds:

$HClO_4$	NaClO	CH_4	$(NH_4)_3PO_4$	CO_3^{2-}
+1 +7 −2	+1 +1 −2	−4 +1	−3 +1 (N, H); +5 −2 (P, O)	+4 −2

NH_4^+

$x + 4\cdot(+1) = +1$

$x = -3$

$x + 3\cdot(-2) = -2$

$X = +4$

Problem 2: Determine if any of the following reactions are redox reactions. If so, identify the substance being oxidized and which is being reduced. Identify the oxidizing agent and the reducing agent.

a. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

−3 +1 0 +4 −2 +1 −2

Oxidation (LEO)

Reduction (GER)

Two groups CH_3 – and CH_2 we treat carbons individually

$CH_3 - CH_2 - CH_3$

$x + 3\cdot(+1) = 0$

$x = -3$

$x + 2\cdot(+1) = 0$

$x = -2$

C_3H_8

$3x + 8\cdot(+1) = 0$

CO_2

$x + 2\cdot(-2) = 0$

$X = +4$

C is the reducing agent.

O_2 is the oxidizing agent.

a) $CH_4(g) + 2 O_2(g) \longrightarrow CO_2(g) + 2 H_2O(g)$

$-4 \quad +1 \qquad\qquad 0 \qquad\qquad +4 \quad -2$

oxidation (LEO)

reduction (GER)

Oxidation — ox. state increases

Reduction — ox. state decreases.

CH_4

$x + 4(+1) = 0$

$x = -4$

CO_2

$x + 2(-2) = 0$

$x = +4$

b. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$0 \quad +1\ -1 \quad +2\ -1 \quad 0$

Oxidation

Reduction

HCl is being reduced, it is the oxidizing agent

Zn is being oxidized, it is the reducing agent

c. $5\ Cr^{3+}(aq) + 3MnO_4^{1-}(aq) + 8\ H_2O(l) \rightarrow 5\ CrO_4^{2-}(aq) + 3Mn^{2+}(aq) + 16H^{1+}$

$+3 \quad +7\ -2 \quad +1\ -2 \quad +6\ -2 \quad +2 \quad +1$

Reduction
MnO_4^- – Oxidizing agent

Oxidation
Cr^{3+} – Reducing agent

MnO_4^-
$x + 4\cdot(-2) = -1$
$x = +7.$

CrO_4^{-2}
$x + 4(-2) = -2$
$x = +6$

d. $Cr_2O_7^{-2}(aq) + 2OH^{-1}(aq) \rightarrow 2CrO_4^{-2}(aq) + H_2O(l)$

$+6\ -2 \quad -2\ +1 \quad +6\ -2 \quad +1\ -2$

No change in the ox. states (Numbers). this is not a redox reaction.

$Cr_2O_7^{-2}$ $\quad 2x + 7\cdot(-2) = -2$
$x = \frac{+12}{2} = +6$

CrO_4^{-2} $\quad x + 4\cdot(-2) = -2$
$x = +6$

e. $2CuCl(aq) \rightarrow CuCl_2(aq) + Cu(s)$

$+1\ -1 \quad +2\ -1 \quad 0$

Oxidation

Reduction

CuCl is being reduced, it is the oxidizing agent
CuCl is being oxidized, it is the reducing agent

$\overset{+}{Cu}\overset{-1}{Cl}$
$+1 \quad -1$

$\overset{+2}{Cu}\overset{-1}{Cl_2}$
$+2 \quad -1$

KEY

GENERAL & ANALYTICAL CHEMISTRY I
CHMG-141
With Dr. Bailey

Name____________

Recitation
Week 11

Part C

Balancing Ox/Red Reactions

Step 1. Write the unbalanced net ionic equation.

↓

Step 2. Decide which atoms are oxidized and which are reduced, and write the two unbalanced half-reactions.

↓

Step 3. Balance both half-reactions for all atoms except O and H.

↓

Step 4. Balance each half-reaction for O by adding water to the side with less O, and balance for H by adding H^+ to the side with less H.

↓

Step 5. Balance each half-reaction for charge by adding electrons to the side with greater positive charge, and then multiply by suitable factors to make the electron count the same in both half-reactions.

↓

Step 6. Add the two balanced half-reactions together, and cancel electrons and other species that appear on both sides of the equation.

↓

Check your answer by making sure the equation is balanced both for atoms and for charge.

Problem A:

Balance the following reaction in (a) acidic solution and (b) basic solution

MnO_2 (s) + H_3AsO_3 (aq) -> Mn^{2+} (aq) + H_3AsO_4 (aq)

+2

(a) acidic solution

1. Split the reaction into two half-reactions.

$$\begin{cases} MnO_2(s) \rightarrow Mn^{2+}(aq) \\ H_3AsO_3(aq) \rightarrow H_3AsO_4(aq) \end{cases}$$

2. Consider the following half reaction:

MnO_2 (s) -> Mn_2+ (aq)

Balance everything but oxygen and hydrogen atoms.

3. Take your answer to (2) and balance the oxygen by adding a water to add oxygen where needed.

$$MnO_2(s) \rightarrow Mn^{2+}(aq) + 2H_2O(l)$$

4. Take your answer to (3) and balance the hydrogen by adding in H^+ where needed (do not forget the charge on the H^+!).

$$MnO_2(s) + 4H^+(aq) \rightarrow Mn^{2+}(aq) + 2H_2O$$

+4 2+

5. Take your answer to (4) and balance the charges by adding electrons to the more positive side of the reaction until the charges are equal.

$$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$$

6. Balance the other half reaction in the same way

H_3AsO_3 (aq) -> H_3AsO_4 (aq)

Add H_2O to balance the O $\quad H_3AsO_3(aq) + \underbrace{H_2O(l)} \rightarrow H_3AsO_4(aq)$

add H^+ to balance the H. $\quad H_3AsO_3(aq) + H_2O(l) \rightarrow H_3AsO_4(aq) + 2H^+$

Balance the charge adding e^- $\rightarrow \quad \overset{0}{H_3AsO_3(aq) + H_2O(l)} \rightarrow \overset{2+}{H_3AsO_4(aq) + 2H^+} + 2e^-$

7. Combine the half reactions to eliminate the electrons

$$H_3AsO_3(aq) + H_2O(l) \rightarrow H_3AsO_4(aq) + 2H^+ + 2e^-$$

$$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$$

$$H_3AsO_3(aq) + \cancel{H_2O}(l) + MnO_2(s) + 4H^+ + \cancel{2e^-} \rightarrow Mn^{2+}(aq) + 2\cancel{H_2O}(l) + H_3AsO_4(aq) + 2H^+ + \cancel{2e^-}$$

8. "Clean up": cancel other species that appear on both sides of the equation.

$$\boxed{H_3AsO_3(aq) + MnO_2(s) + 2H^+(aq) \rightarrow Mn^{2+}(aq) + H_2O(l) + H_3AsO_4(aq)}$$

(b) Basic solution

Add OH- to both sides of the equation to neutralize H+

$$H_3AsO_3(aq) + MnO_2(s) + \overbrace{2H^+(aq) + 2OH^-(aq)}^{=2H_2O(l)} \longrightarrow Mn^{2+}(aq) + \cancel{H_2O(l)} + H_3AsO_4(aq) + 2OH^-(aq)$$

$$\boxed{H_3AsO_3(aq) + MnO_2(s) + H_2O(l) \longrightarrow Mn^{2+}(aq) + H_3AsO_4(aq) + 2OH^-(aq)}$$

<u>Problem B (additional, for your practice):</u>

Balance the following reaction in basic solution

$CrO_4{}^{2-}$ (aq) + I_2 (s) -> $Cr(OH)_3$ (s) + $IO_3{}^-$ (aq)

Half reactions:

$$\begin{cases} CrO_4^{2-}(aq) \longrightarrow Cr(OH)_3(s) \\ I_2(s) \longrightarrow IO_3^-(aq) \end{cases}$$

CrO_4^{2-} (aq) + I_2 (s) → $Cr(OH)_3$ (s) + IO_3^- (aq)

1. Separate into half reactions:

CrO_4^{2-} (aq) → $Cr(OH)_3$ (s) GER reduction
$+6$ $+3$

I_2 (s) → IO_3^- (aq) LEO oxidation
0 $+5$

CrO_4^{2-}
$x + 4(-2) = -2$
$x = 8 - 2 = +6$

$Cr(OH)_3$
$x + 3(-2) + 3(+1) = 0$
$x = +3$

IO_3^-
$x + 3(-2) = -1$
$x = +5$

See also step 5

2. Balance everything but H and O

CrO_4^{2-} (aq) → $Cr(OH)_3$ (s)

I_2 (s) → 2 IO_3^- (aq)

3. Balance O with H_2O

CrO_4^{2-} (aq) → $Cr(OH)_3$ (s) + H_2O (l)

I_2 (s) + 6 H_2O (l) → 2 IO_3^- (aq)

4. Balance H with H+

CrO_4^{2-} (aq) + 5 H^+ → $Cr(OH)_3$ (s) + H_2O (l)

I_2 (s) + 6 H_2O (l) → 2 IO_3^- (aq) + 12 H^+

5. Balance charge with e-

CrO_4^{2-} (aq) + 5 H^+ + 3 e- $\rightarrow$ $Cr(OH)_3$ (s) + H_2O (l)
+6 +3

I_2 (s) + 6 H_2O (l) $\rightarrow$ 2 IO_3^- (aq) + 12 H^+ + 10e-
0 +5

GER

LEO

6. Combine to get rid of e-

10x[CrO_4^{2-} (aq) + 5 H^+ + 3 e- $\rightarrow$ $Cr(OH)_3$ (s) + H_2O (l)]

3x[I_2 (s) + 6 H_2O (l) $\rightarrow$ 2 IO_3^- (aq) + 12 H^+ + 10e-]

10 CrO_4^{2-} (aq) + 50 H^+ + 30 e- + 3 I_2 (s) + 18 H_2O (l) $\rightarrow$ 10 $Cr(OH)_3$ (s) +10 H_2O (l)+ 6 IO_3^- (aq) + 36 H^+ + 30 e-

Let's clean up a little bit:

10 CrO_4^{2-} (aq) + 14 ~~50~~ H^+ + ~~30 e-~~ + 3 I_2 (s) + 8 ~~18~~ H_2O (l) $\rightarrow$ 10 $Cr(OH)_3$ (s) + ~~10 H_2O (l)~~ + 6 IO_3^- (aq) + ~~36 H^+ + 30 e-~~

In acidic solution

$\boxed{10\ CrO_4^{2-}\ (aq) + 14\ H^+ + 3\ I_2\ (s) + 8\ H_2O\ (l) \rightarrow 10\ Cr(OH)_3\ (s) + 6\ IO_3^-\ (aq)}$

In Basic solution:

7. Add OH- to neutralize H+ to both sides of the reaction:

10 CrO_4^{2-} (aq) + 14 H^+ + 3 I_2 (s) + 8 H_2O (l) + 14 OH- $\rightarrow$ 10 $Cr(OH)_3$ (s) + 6 IO_3^- (aq) + 14 OH-
~~OH-~~

10 CrO_4^{2-} (aq) + 14 H_2O + 3 I_2 (s) + 8 H_2O (l) - $\rightarrow$ 10 $Cr(OH)_3$ (s) + 6 IO_3^- (aq) + 14 OH-

$\boxed{10\ CrO_4^{2-}\ (aq) + 22\ H_2O + 3\ I_2\ (s) \rightarrow 10\ Cr(OH)_3\ (s) + 6\ IO_3^-\ (aq) + 14\ OH^-}$