

Name: \_\_\_\_\_

Period: \_\_\_\_\_

Seat#: \_\_\_\_\_

**Directions:** Use the rules for Assigning Oxidation numbers to determine the oxidation number assigned to each element in each of the given formulas. The rules are at the bottom of the page.

Q #	Formula	Oxidation #'s	
1)	Cl <sub>2</sub>	Cl	
2)	Cl <sup>-</sup>	Cl	
3)	Na	Na	
4)	Na <sup>+</sup>	Na	
5)	KCl	K	Cl
6)	H <sub>2</sub> S	H	S
7)	CaO	Ca	O
8)	H <sub>2</sub> O	H	O
9)	NO <sub>3</sub> <sup>-</sup>	N	O
10)	NO <sub>2</sub>	N	O
11)	Cr <sub>2</sub> O <sub>7</sub> <sup>2-</sup>	Cr	O
12)	O <sub>2</sub>	O	
13)	NH <sub>3</sub>	N	H
14)	CaH <sub>2</sub>	Ca	H
15)	SO <sub>4</sub> <sup>2-</sup>	S	O

Q #	Formula	Oxidation #'s		
16)	Na <sub>2</sub> O	Na	O	
17)	HNO <sub>3</sub>	H	N	O
18)	CaCl <sub>2</sub>	Ca	Cl	
19)	PO <sub>4</sub> <sup>3-</sup>	P	O	
20)	MnO <sub>2</sub>	Mn	O	
21)	K <sub>3</sub> PO <sub>4</sub>	K	P	O
22)	Fe <sub>2</sub> O <sub>3</sub>	Fe	O	
23)	KNO <sub>2</sub>	K	N	O
24)	N <sub>2</sub>	N		
25)	Al <sup>3+</sup>	Al		
26)	H <sub>2</sub> O <sub>2</sub>	H	O	
27)	H <sub>2</sub> SO <sub>4</sub>	H	S	O
28)	NH <sub>4</sub> Cl	N	H	Cl
29)	FeO	Fe	O	
30)	SiO <sub>2</sub>	Si	O	

### Rules for Assigning Oxidation Numbers

- 1) The oxidation number of any uncombined element is 0.
- 2) The oxidation number of a monatomic ion equals the charge on the ion.
- 3) The more-electronegative element in a binary compound is assigned the number equal to the charge it would have if it were an ion.
- 4) The oxidation number of fluorine in a compound is always -1
- 5) Oxygen has an oxidation number of -2 unless it is combined with F, when it is +2, or it is in a peroxide, such as H<sub>2</sub>O<sub>2</sub>, when it is -1
- 6) The oxidation state of H in most of its compounds is +1 unless it is combined with a metal, in which case it is -1
- 7) In compounds, the elements of groups 1 and 2 as well as aluminum have oxidation numbers +1, +2 and +3 respectively.
- 8) The sum of the oxidation numbers of all atoms in a neutral compound is 0.
- 9) The sum of the oxidation numbers of all atoms in a polyatomic ion equals charge of the ion.



Name: \_\_\_\_\_

Period: \_\_\_\_\_

Seat#: \_\_\_\_\_

**Directions:** In the left hand column, identify the oxidation states of the elements undergoing reduction/oxidation. In the right hand column balance the half reactions.

**Rules:**

- |  |   |
|--|---|
| <ol style="list-style-type: none"> <li>1. Find oxidation #'s</li> <li>2. Determine which elements are reduced/oxidized</li> <li>3. Write each half reaction separately</li> <li>4. Balance "unique" atoms<br/>(everything except oxygens and hydrogens)</li> </ol> | <ol style="list-style-type: none"> <li>5. Add H<sub>2</sub>O's to balance any oxygens</li> <li>6. Add H<sup>+</sup>'s to balance any hydrogens</li> <li>7. Add e<sup>-</sup>'s to balance the charge</li> </ol> |
|--|---|

<b>1)</b> $\text{Li} + \text{F}_2 \rightarrow 2\text{F}^- + \text{Li}^+$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction
<b>2)</b> $\text{Pb}^{2+} + \text{Mn}^{2+} \rightarrow \text{MnO}_2 + \text{Pb}$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction
<b>3)</b> $\text{Cl}_2 + 2\text{Br}^- \rightarrow 2\text{Cl}^- + \text{Br}_2$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction
<b>4)</b> $\text{Mg} + \text{NO}_3^- \rightarrow \text{Mg}^{2+} + \text{NO}$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction

**Dougherty Valley HS Chemistry**  
**Redox – Writing Half Reactions**

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<b>5)</b> $\text{MnO}_4^- + \text{Pb} \rightarrow \text{Pb}^{2+} + \text{Mn}^{2+}$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction
<b>6)</b> $\text{Fe}_2\text{O}_3(\text{s}) + 2\text{Al}(\text{s}) \rightarrow 2\text{Fe}(\text{l}) + \text{Al}_2\text{O}_3(\text{s})$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction
<b>7)</b> $2\text{Ag} + \text{Ce}^{4+} \rightarrow \text{Ag}_2\text{O}_2 + \text{Ce}^{3+}$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction
<b>8)</b> $\text{PbO}_2 + \text{Ag} \rightarrow \text{Ag}^+ + \text{Pb}^{2+}$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction
<b>9)</b> $\text{Hg}_2^{2+} + \text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{Hg}$	Balance Oxidation Half Reaction
	Balance Reduction Half Reaction

Name: \_\_\_\_\_

Period: \_\_\_\_\_

Seat#: \_\_\_\_\_

**Directions:** In each of the following chemical compounds, determine the oxidation states of each element.

Q#	Name	Formula	Oxidation State for Each Element
1)	Sodium nitrate		
2)	Ammonia		
3)	Zinc oxide		
4)	Water		
5)	Calcium hydride		
6)	Carbon dioxide		
7)	Nitrogen		
8)	Sodium sulfate		
9)	Aluminum hydroxide		
10)	Magnesium phosphate		

**Directions:** In each of the following, determine what was oxidized and what was reduced.

Q#	Reaction	Element Oxidized	Element Reduced
11)	$\text{Ca} + \text{H}_2\text{O} \rightarrow \text{CaO} + \text{H}_2$		
12)	$2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O}$		



Name: \_\_\_\_\_

Period: \_\_\_\_\_

Seat#: \_\_\_\_\_

**Directions:** Answer the questions below. Remember that individual elements are oxidized/reduced, but the reactant molecule as a whole is the “oxidizing agent” or the “reducing agent.”

1) What is at least one mnemonic (or other memory device) you can use to help you with Redox topics?

2) Consider the following redox reaction:  $\text{KMnO}_4(\text{aq}) + \text{NO}(\text{g}) \rightarrow \text{MnO}_2(\text{s}) + \text{KNO}_3(\text{aq})$

a. Give the **oxidation number** for each element in the compounds below:

$\text{KMnO}_4$ : K = \_\_\_\_\_, Mn = \_\_\_\_\_, O = \_\_\_\_\_       $\text{MnO}_2$ : Mn = \_\_\_\_\_, O = \_\_\_\_\_

$\text{NO}$ : N = \_\_\_\_\_, O = \_\_\_\_\_       $\text{KNO}_3$ : K = \_\_\_\_\_, N = \_\_\_\_\_, O = \_\_\_\_\_

b. The element oxidized is \_\_\_\_\_, and the oxidizing agent is \_\_\_\_\_.

The element reduced is \_\_\_\_\_, and the reducing agent is \_\_\_\_\_.

c. The total number of electrons transferred in this reaction is \_\_\_\_\_.

3) Consider the following redox reaction:  $\text{Fe}_2\text{O}_3(\text{s}) + 3 \text{CO}(\text{g}) \rightarrow 2 \text{Fe}(\text{s}) + 3 \text{CO}_2(\text{g})$

a. Give the **oxidation number** for each element in the compounds below:

$\text{Fe}_2\text{O}_3$ : Fe = \_\_\_\_\_, O = \_\_\_\_\_      Fe: Fe = \_\_\_\_\_

$\text{CO}$ : C = \_\_\_\_\_, O = \_\_\_\_\_       $\text{CO}_2$ : C = \_\_\_\_\_, O = \_\_\_\_\_

b. The element oxidized is \_\_\_\_\_, and the oxidizing agent is \_\_\_\_\_.

The element reduced is \_\_\_\_\_, and the reducing agent is \_\_\_\_\_.

c. The total number of electrons transferred in this reaction is \_\_\_\_\_.

4) Consider the following :  $\text{H}_2\text{SO}_3(\text{aq}) + 2 \text{Mn}(\text{s}) + 4 \text{HCl}(\text{aq}) \rightarrow \text{S}(\text{s}) + 2 \text{MnCl}_2(\text{aq}) + 3 \text{H}_2\text{O}(\text{l})$

a. Give the **oxidation number** for each element in the compounds below:

$\text{H}_2\text{SO}_3$ : H = \_\_\_\_\_, S = \_\_\_\_\_, O = \_\_\_\_\_      S: S = \_\_\_\_\_

$\text{Mn}$ : Mn = \_\_\_\_\_       $\text{MnCl}_2$ : Mn = \_\_\_\_\_, Cl = \_\_\_\_\_

$\text{HCl}$ : H = \_\_\_\_\_, Cl = \_\_\_\_\_       $\text{H}_2\text{O}$ : H = \_\_\_\_\_, O = \_\_\_\_\_

b. The element oxidized is \_\_\_\_\_, and the oxidizing agent is \_\_\_\_\_.

The element reduced is \_\_\_\_\_, and the reducing agent is \_\_\_\_\_.

c. The total number of electrons transferred in this reaction is \_\_\_\_\_.



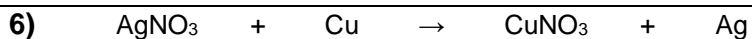
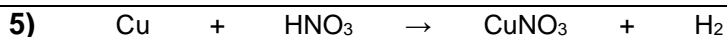
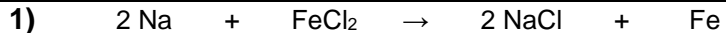


Name: \_\_\_\_\_

Period: \_\_\_\_\_

Seat#: \_\_\_\_\_

**Directions:** In each of the following equations, indicate the element that has been oxidized and the one that has been reduced. You should also label the oxidation state of each before and after the process. Also identify the oxidizing and reducing agent.





Name: \_\_\_\_\_

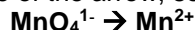
Period: \_\_\_\_\_

Seat#: \_\_\_\_\_

**For Half-Reactions in Acidic Solution**

**Step One: Balance the atom being reduced/oxidized.**

In our example, there is already one Mn on each side of the arrow, so this step is already done.



**Step Two: Balance the oxygen's.**

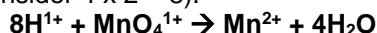
Do this by adding water molecules (as many as are needed) to the side needing oxygen. In our case, the left side has 4 oxygen's, while the right side has none, so:



Notice that, when the water is added, hydrogen's also come along. There is nothing that can be done about this; we'll take care of it in the next step. A common question is: "Why can't I just add 4 oxygen atoms to the right side?" Quick answer: don't do it, it's wrong. There are not free oxygen atoms floating around in our reaction vessel. So you can't add them!

**Step Three: Balance the hydrogen's.**

Do this by adding hydrogen ions (as many as are needed) to the side needing hydrogen. In our example, we need 8 (notice the water molecule's formula, then consider  $4 \times 2 = 8$ ).

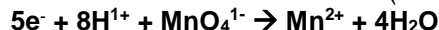


**Step Four: Balance the total charge.**

This will be done using electrons. It is ALWAYS the last step. First, a comment. You do not need to look at the oxidation number for each atom. You only need to look at the charge on the ion or molecule, then sum those up.

Left side of the reaction, total charge is +7. There are 8 H+, giving  $8 \times +1 = +8$  and a minus one from the permanganate. (A very typical wrong answer for the left side is zero. Someone only counted the +1 and the -1, they forget the 8.)

Right side of the reaction, total charge is +2. The water molecule is neutral (zero charge) and the single Mn is +2.



Five electrons reduces the +7 to a +2 and the two sides are EQUAL in total charge.

The half-reaction is now correctly balanced!

**For Half-Reactions in Basic Solution**

**Step One to Four: Balance the half-reaction AS IF it were in acid solution.**

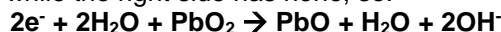
Just go about steps 1-4 business as usual! The half-reaction is actually in basic solution, but we are going to start out as if it were in acid solution and then fix it later to account for it being in a basic solution. Here are the 4 acid steps:

- 1) Balance the atom being reduced/oxidized.
- 2) Balance the oxygens.
- 3) Balance the hydrogens.
- 4) Balance the charge.

Let's use this half reaction for our example. Pretend we just finished steps 1-4:  $2e^- + 2\text{H}^{1+} + \text{PbO}_2 \rightarrow \text{PbO} + \text{H}_2\text{O}$

**Step Five: Convert all H<sup>+</sup> to H<sub>2</sub>O.**

If you are in a basic solution you will not have H<sup>+</sup> floating around. It is combined with the hydroxides that are present from the base! To convert the H<sup>+</sup> into water add the same number of OH<sup>-</sup> ions as you have H<sup>+</sup> ions, BUT you have to add them to both sides to preserve the balancing! The side with the H<sup>+</sup> will determine how many hydroxide to add. In our case, the left side has 2 hydrogen ions, while the right side has none, so:



Notice that, when the two hydroxide ions on the left were added, they immediately reacted with the hydrogen ion present. The reaction is:  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

**Step Six: Remove any duplicate molecules or ions.**

In our example, there are two water molecules on the left and one on the right. This means one water molecule may be removed from each side, giving:



The half-reaction is now correctly balanced!

*By the way, notice the 2OH<sup>-</sup>. Be careful to read that as two hydroxide ions (2 OH<sup>-</sup>) and NOT twenty hydride ions (20 H<sup>-</sup>). People have been known to do that.*

## Dougherty Valley HS Chemistry

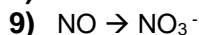
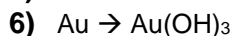
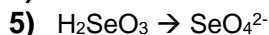
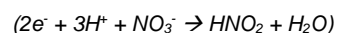
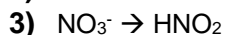
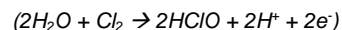
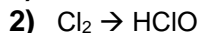
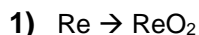
### Redox – Balancing Redox Rxns in Acidic and Basic Solutions

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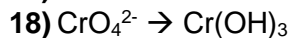
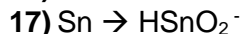
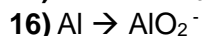
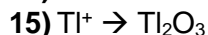
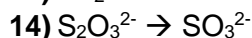
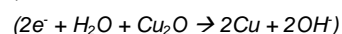
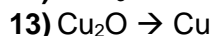
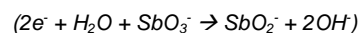
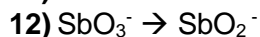
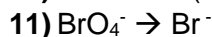
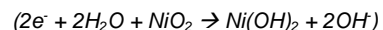
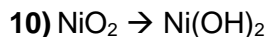
**Directions:** In each of the following equations do the following on binder paper:

- Balance the reaction in the type of solution indicated.
- If you need/want some extra practice:
  - Indicate the element that has been oxidized and the one that has been reduced.
  - Label the oxidation state of each before and after the process.
  - Some answers have been provided in italics to the side.

#### Reaction done in Acidic Solution



#### Reaction done in Basic Solution



Name: \_\_\_\_\_

Period: \_\_\_\_\_

Seat#: \_\_\_\_\_

**Directions:** Any worksheet that is labeled with an \* means it is suggested extra practice. We do not always have time to assign every possible worksheet that would be good practice for you to do. You can do this worksheet when you have extra time, when you finish something early, or to help you study for a quiz or a test. If and when you choose to do this Extra Practice worksheet, please do the work on binder paper. You will include this paper stapled into your Rainbow Packet when you turn it in, even if you didn't do any of this. We want to make sure we keep it where it belongs so you can do it later if you want to (or need to). If you did the work on binder paper you can include that in your Rainbow Packet after this worksheet. If we end up with extra class time then portions of this may turn into required work. If that happens you will be told which problems are turned into required. Remember there is tons of other extra practice on the class website...and the entire internet! See me if you need help finding practice on a topic you are struggling with.

**\*Note\*** Some of these are really nasty...redox equations get awful really fast! Just go slow, use lots of binder paper, show each step, double check your counting, if you get really stuck then erase and start over.

**Balance these reactions. Some are in acidic solution, some are in basic solution. Some answers are provided in italics. I do NOT have a full worked out key for these, or answers to the ones without answers provided.**

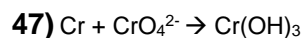
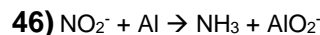
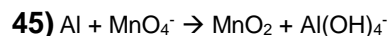
### Acidic Solution

- 1)  $\text{ClO}_3^- + \text{SO}_2 \rightarrow \text{SO}_4^{2-} + \text{Cl}^-$   
 *$\text{H}_2\text{O} + \text{ClO}_3^- + 3\text{SO}_2 \rightarrow 3\text{SO}_4^{2-} + \text{Cl}^- + 6\text{H}^+$*
- 2)  $\text{H}_2\text{S} + \text{NO}_3^- \rightarrow \text{S}_8 + \text{NO}$   
 *$16\text{H}^+ + 24\text{H}_2\text{S} + 16\text{NO}_3^- \rightarrow 3\text{S}_8 + 16\text{NO} + 32\text{H}_2\text{O}$*
- 3)  $\text{MnO}_4^- + \text{H}_2\text{S} \rightarrow \text{Mn}^{2+} + \text{S}_8$   
 *$48\text{H}^+ + 16\text{MnO}_4^- + 40\text{H}_2\text{S} \rightarrow 16\text{Mn}^{2+} + 5\text{S}_8 + 64\text{H}_2\text{O}$*
- 4)  $\text{Cu} + \text{SO}_4^{2-} \rightarrow \text{Cu}^{2+} + \text{SO}_2$   
 *$4\text{H}^+ + \text{Cu} + \text{SO}_4^{2-} \rightarrow \text{Cu}^{2+} + \text{SO}_2 + 2\text{H}_2\text{O}$*
- 5)  $\text{MnO}_4^- + \text{CH}_3\text{OH} \rightarrow \text{CH}_3\text{COOH} + \text{Mn}^{2+}$
- 6)  $\text{Cr}_2\text{O}_7^{2-} + \text{Fe}^{2+} \rightarrow \text{Cr}^{3+} + \text{Fe}^{3+}$
- 7)  $\text{HNO}_2 \rightarrow \text{NO} + \text{NO}_2$
- 8)  $\text{H}_2\text{C}_2\text{O}_4 + \text{MnO}_4^- \rightarrow \text{CO}_2 + \text{Mn}^{2+}$
- 9)  $\text{O}_2 + \text{As} \rightarrow \text{HASO}_2 + \text{H}_2\text{O}$
- 10)  $\text{NO}_2 \rightarrow \text{NO}_3^- + \text{NO}$
- 11)  $\text{ClO}_4^- + \text{Cl}^- \rightarrow \text{ClO}^- + \text{Cl}_2$
- 12)  $\text{H}_5\text{IO}_6 + \text{Cr} \rightarrow \text{IO}_3^- + \text{Cr}^{3+}$
- 13)  $\text{Fe} + \text{HCl} \rightarrow \text{HFeCl}_4 + \text{H}_2$
- 14)  $\text{NO}_3^- + \text{H}_2\text{O}_2 \rightarrow \text{NO} + \text{O}_2$
- 15)  $\text{BrO}_3^- + \text{Fe}^{2+} \rightarrow \text{Br}^- + \text{Fe}^{3+}$
- 16)  $\text{Cr}_2\text{O}_7^{2-} + \text{C}_2\text{H}_4\text{O} \rightarrow \text{CH}_3\text{COOH} + \text{Cr}^{3+}$
- 17)  $\text{MnO}_4^- + \text{C}_2\text{H}_4\text{O} \rightarrow \text{CH}_3\text{COOH} + \text{MnO}_2$
- 18)  $\text{Zn} + \text{NO}_3^- \rightarrow \text{NH}_4^+ + \text{Zn}^{2+}$
- 19)  $\text{HBr} + \text{SO}_4^{2-} \rightarrow \text{SO}_2 + \text{Br}_2$
- 20)  $\text{NO}_3^- + \text{I}_2 \rightarrow \text{IO}_3^- + \text{NO}_2$
- 21)  $\text{CuS} + \text{NO}_3^- \rightarrow \text{NO} + \text{Cu}^{2+} + \text{HSO}_4^-$
- 22)  $\text{I}^- + \text{ClO}^- \rightarrow \text{I}_3^- + \text{Cl}^-$

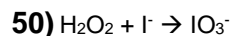
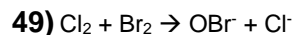
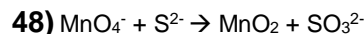
- 23)  $\text{As}_2\text{O}_3 + \text{NO}_3^- \rightarrow \text{H}_3\text{AsO}_4 + \text{NO}$
- 24)  $\text{Br}^- + \text{MnO}_4^- \rightarrow \text{Br}_2 + \text{Mn}^{2+}$
- 25)  $\text{CH}_3\text{OH} + \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{CH}_2\text{O} + \text{Cr}^{3+}$
- 26)  $\text{Mn}^{2+} + \text{BiO}_3^- \rightarrow \text{Bi}^{3+} + \text{MnO}_4^-$
- 27)  $\text{S}_8 + \text{NO}_3^- \rightarrow \text{SO}_3^{2-} + \text{NO}$
- 28)  $\text{H}_3\text{AsO}_4 + \text{Zn} \rightarrow \text{AsH}_3 + \text{Zn}^{2+}$
- 29)  $\text{P}_4 + \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{H}_3\text{PO}_4 + \text{Cr}^{3+}$

### Basic Solution

- 30)  $\text{NH}_3 + \text{ClO}^- \rightarrow \text{N}_2\text{H}_4 + \text{Cl}^-$   
 *$2\text{NH}_3 + \text{ClO}^- \rightarrow \text{N}_2\text{H}_4 + \text{Cl}^- + \text{H}_2\text{O}$*
- 31)  $\text{Au} + \text{O}_2 + \text{CN}^- \rightarrow \text{Au}(\text{CN})_2^- + \text{H}_2\text{O}_2$   
 *$2\text{H}_2\text{O} + 2\text{Au} + \text{O}_2 + 4\text{CN}^- \rightarrow 2\text{Au}(\text{CN})_2^- + \text{H}_2\text{O}_2 + 2\text{OH}^-$*
- 32)  $\text{AlH}_4^- + \text{H}_2\text{CO} \rightarrow \text{Al}^{3+} + \text{CH}_3\text{OH}$   
 *$4\text{H}_2\text{O} + \text{AlH}_4^- + 4\text{H}_2\text{CO} \rightarrow \text{Al}^{3+} + 4\text{CH}_3\text{OH} + 4\text{OH}^-$*
- 33)  $\text{Se} + \text{Cr}(\text{OH})_3 \rightarrow \text{Cr} + \text{SeO}_3^{2-}$   
 *$6\text{OH}^- + 3\text{Se} + 4\text{Cr}(\text{OH})_3 \rightarrow 4\text{Cr} + 3\text{SeO}_3^{2-} + 9\text{H}_2\text{O}$*
- 34)  $\text{Br}^- + \text{MnO}_4^- \rightarrow \text{MnO}_2 + \text{BrO}_3^-$
- 35)  $\text{H}_2\text{O}_2 + \text{Cl}_2\text{O}_7 \rightarrow \text{ClO}_2^- + \text{O}_2$
- 36)  $\text{Fe} + \text{NiO}_2 \rightarrow \text{Fe}(\text{OH})_2 + \text{Ni}(\text{OH})_2$
- 37)  $\text{MnO}_4^- + \text{H}_2\text{O}_2 \rightarrow \text{MnO}_2 + \text{O}_2$
- 38)  $\text{Zn} + \text{BrO}_4^- \rightarrow [\text{Zn}(\text{OH})_4]^{2-} + \text{Br}^-$
- 39)  $\text{MnO}_4^- + \text{S}^{2-} \rightarrow \text{MnO}_2 + \text{S}_8$
- 40)  $\text{Pb}(\text{OH})_4^{2-} + \text{ClO}^- \rightarrow \text{PbO}_2 + \text{Cl}^-$
- 41)  $\text{Ti}_2\text{O}_3 + \text{NH}_2\text{OH} \rightarrow \text{TiOH} + \text{N}_2$
- 42)  $\text{Fe}(\text{OH})_2 + \text{CrO}_4^{2-} \rightarrow \text{Fe}_2\text{O}_3 + \text{Cr}(\text{OH})_4^-$
- 43)  $\text{Br}_2 \rightarrow \text{Br}^- + \text{BrO}_3^-$
- 44)  $\text{ClO}_2^- + \text{H}_2\text{O} \rightarrow \text{ClO}_2 + \text{OH}^-$



*Note: Cr(OH)<sub>3</sub> is found in BOTH half reactions!*



*Note: IO<sub>3</sub><sup>-</sup> is found in both half reactions!*

