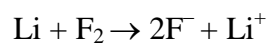


Directions: Use the rules for Assigning Oxidation numbers to determine the oxidation number assigned to each element in each of the given formulas.

	Formula				Formula			
1.	Cl ₂	Cl		16.	Na ₂ O ₂	Na	O	
2.	Cl ⁻	Cl		17.	HNO ₃	H	N	O
3.	Na	Na		18.	CaCl ₂	Ca	Cl	
4.	Na ⁺	Na		19.	PO ₄ ³⁻	P	O	
5.	KCl	K	Cl	20.	MnO ₂	Mn	O	
6.	H ₂ S	H	S	21.	K ₃ PO ₄	K	P	O
7.	CaO	Ca	O	22.	Fe ₂ O ₃	Fe	O	
8.	H ₂ O	H	O	23.	KNO ₂	K	N	O
9.	NO ₃ ⁻	N	O	24.	N ₂	N		
10.	NO ₂	N	O	25.	Al ³⁺	Al		
11.	Cr ₂ O ₇ ²⁻	Cr	O	26.	H ₂ O ₂	H	O	
12.	O ₂	O		27.	H ₂ SO ₄	H	S	O
13.	NH ₃	N	H	28.	NH ₄ Cl	N	H	Cl
14.	CaH ₂	Ca	H	29.	FeO	Fe	O	
15.	SO ₄ ²⁻	S	O	30.	SiO ₂	Si	O	

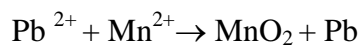
Rules for Assigning Oxidation Numbers

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 H₂O₂, when it is -1.
6. The oxidation state of hydrogen 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.



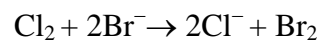
Balance Oxidation half reaction:

Balance Reduction half reaction:



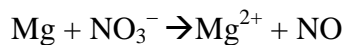
Balance Oxidation half reaction:

Balance Reduction half reaction:



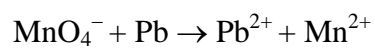
Balance Oxidation half reaction:

Balance Reduction half reaction:



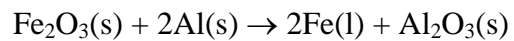
Balance Oxidation half reaction:

Balance Reduction half reaction:



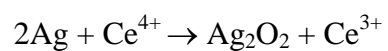
Balance Oxidation half reaction:

Balance Reduction half reaction:



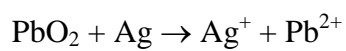
Balance Oxidation half reaction:

Balance Reduction half reaction:



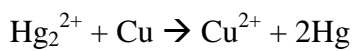
Balance Oxidation half reaction:

Balance Reduction half reaction:



Balance Oxidation half reaction:

Balance Reduction half reaction:



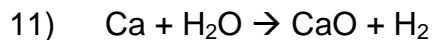
Balance Oxidation half reaction:

Balance Reduction half reaction:

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

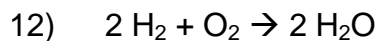
- 1) Sodium nitrate _____
- 2) Ammonia _____
- 3) Zinc oxide _____
- 4) Water _____
- 5) Calcium hydride _____
- 6) Carbon dioxide _____
- 7) Nitrogen _____
- 8) Sodium sulfate _____
- 9) Aluminum hydroxide _____
- 10) Magnesium phosphate _____

In each of the following reactions, determine what was oxidized and what was reduced.



Element oxidized: _____

Element reduced: _____



Element oxidized: _____

Element reduced: _____

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

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

In KMnO_4 : K = _____, Mn = _____, and O = _____.

In KNO_3 : K = _____, N = _____, and O = _____.

Indicate the *neutral element* or the *ion (or element with its oxidation number) in a compound* that is oxidized or reduced and the *reactants* that served as the oxidizing and the reducing agents below:

b. The reactant oxidized is _____, and the oxidizing agent is _____.

The reactant reduced is _____, and the reducing agent is _____.

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

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

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

In CO_2 : C = _____, and O = _____.

b. The reactant oxidized is _____, and the oxidizing agent is _____.

The reactant reduced is _____, and the reducing agent is _____.

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

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

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

In H_2SO_3 : H = _____, S = _____, and O = _____.

In MnCl_2 : Mn = _____ and Cl = _____.

b. The reactant oxidized is _____, and the oxidizing agent is _____.

The reactant reduced is _____, and the reducing agent is _____.

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

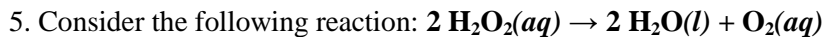


Indicate the *neutral element* or the *ion (or element with its oxidation number) in a compound* that is oxidized or reduced and the *reactants* that served as the oxidizing and the reducing agents below:

a. The reactant oxidized is _____, and the oxidizing agent is _____.

The reactant reduced is _____, and the reducing agent is _____.

b. The total number of electrons transferred in this reaction is _____.



Indicate the *neutral element* or the *ion (or element with its oxidation number) in a compound* that is oxidized or reduced and the *reactants* that served as the oxidizing and the reducing agents below:

a. The reactant oxidized is _____, and the oxidizing agent is _____.

The reactant reduced is _____, and the reducing agent is _____.

b. The total number of electrons transferred in this reaction is _____.



Indicate the *neutral element* or the *ion (or element with its oxidation number) in a compound* that is oxidized or reduced and the *reactants* that served as the oxidizing and the reducing agents below:

a. The reactant oxidized is _____, and the oxidizing agent is _____.

The reactant reduced is _____, and the reducing agent is _____.

b. The total number of electrons transferred in this reaction is _____.



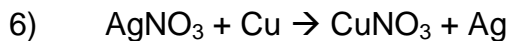
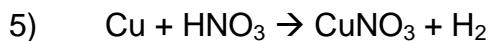
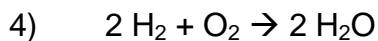
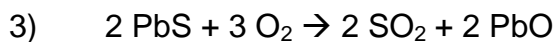
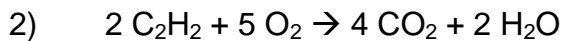
Indicate the *neutral element* or the *ion (or element with its oxidation number) in a compound* that is oxidized or reduced and the *reactants* that served as the oxidizing and the reducing agents below:

a. The reactant oxidized is _____, and the oxidizing agent is _____.

The reactant reduced is _____, and the reducing agent is _____.

b. The total number of electrons transferred in this reaction is _____.

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:



Complete the following reactions on a separate sheet of paper, showing all steps involved, neatly and clearly.

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.

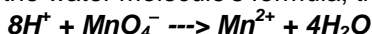


Step Two: Balance the oxygen's. Do this by adding water molecules (as many as are needed) to the side needing oxygen. In our case, the left side has 4 oxygen's, while the right side has none, so:



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. The "why" will be left to another day.

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 $4 \times 2 = 8$).

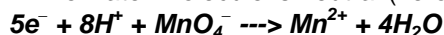


Step Four: Balance the total charge. This will be done using electrons. It is ALWAYS the last step.

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. The person sees only the +1 and the -1, they forget the 8. When you do this step in the parallel example, don't forget to multiply 2 times 3. I'll leave you to figure out where in the problem that is.)

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



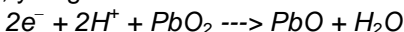
Five electrons reduces the +7 to a +2 and the two sides are EQUAL in total charge. The half-reaction is now correctly balanced.

For Half-Reactions in Basic Solution

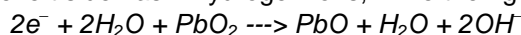
Step One to Four: Balance the half-reaction AS IF it were in acid solution. I hope you got that. The half-reaction is actually in basic solution, but we are going to start out as if it were in acid solution. Here are the 4 acid steps:

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

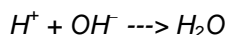
When you do that to the above half-reaction, you get:



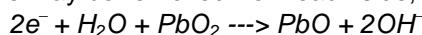
Step Five: Convert all H^+ to H_2O . Do this by adding OH^- ions to both sides. The side with the H^+ will determine how many hydroxide to add. In our case, the left side has 2 hydrogen ions, while the right side has none, so:



Notice that, when the two hydroxide ions on the left were added, they immediately reacted with the hydrogen ion present. The reaction is:



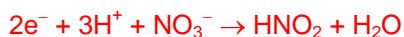
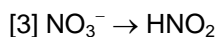
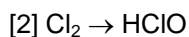
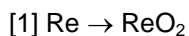
Step Six: Remove any duplicate molecules or ions. In our example, there are two water molecules on the left and one on the right. This means one water molecule may be removed from each side, giving:

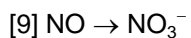
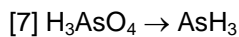
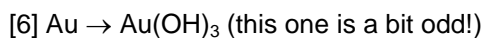
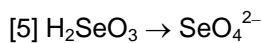


The half-reaction is now correctly balanced.

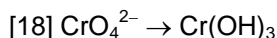
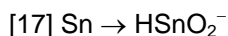
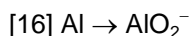
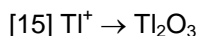
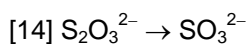
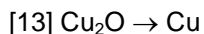
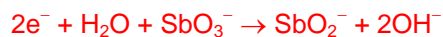
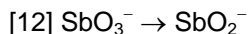
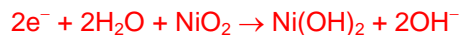
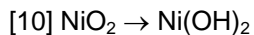
By the way, notice the 2OH^- . Be careful to read that as two hydroxide ions (2OH^-) and NOT twenty hydride ions (20H^-). People have been known to do that.

Balance each half-reaction, the reaction being in acidic solution.

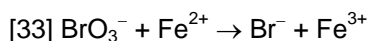
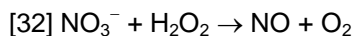
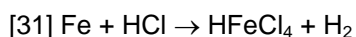
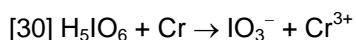
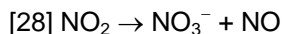
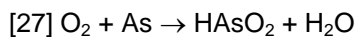
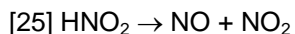
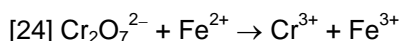
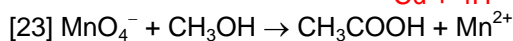
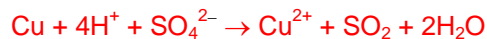
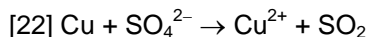
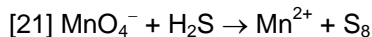
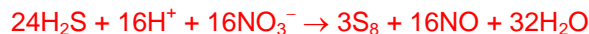
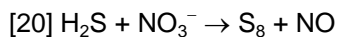
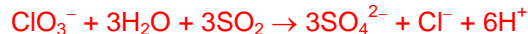
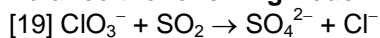


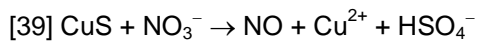
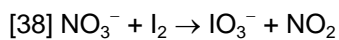
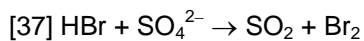
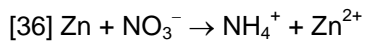
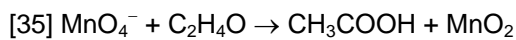
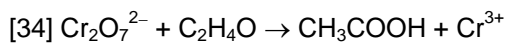


Balance each half-reaction, the reaction being in basic solution.

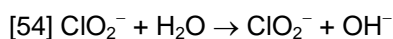
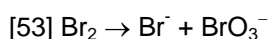
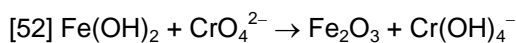
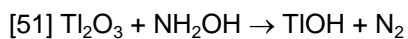
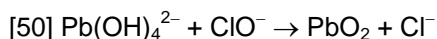
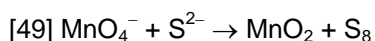
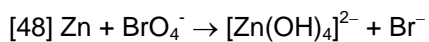
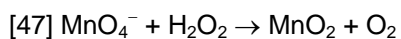
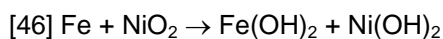
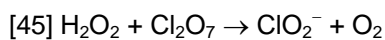
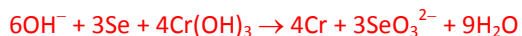
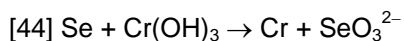
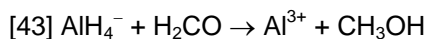
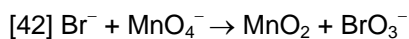
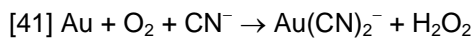


Balance the following Redox reactions in Acidic Solution





Balance the following Redox Reactions in Basic Solution



Redox Reactions in Acidic Solution:

1. $\Gamma^{-}(\text{aq}) + \text{ClO}^{-}(\text{aq}) \longrightarrow \text{I}_3^{-}(\text{aq}) + \text{Cl}^{-}(\text{aq})$
2. $\text{As}_2\text{O}_3(\text{s}) + \text{NO}_3^{-}(\text{aq}) \longrightarrow \text{H}_3\text{AsO}_4(\text{aq}) + \text{NO}(\text{g})$
3. $\text{Br}^{-}(\text{aq}) + \text{MnO}_4^{-}(\text{aq}) \longrightarrow \text{Br}_2(\text{l}) + \text{Mn}^{2+}(\text{aq})$
4. $\text{CH}_3\text{OH}(\text{aq}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) \longrightarrow \text{CH}_2\text{O}(\text{l}) + \text{Cr}^{3+}(\text{aq})$
5. $\text{Mn}^{2+}(\text{aq}) + \text{BiO}_3^{-}(\text{aq}) \longrightarrow \text{Bi}^{3+}(\text{aq}) + \text{MnO}_4^{-}(\text{aq})$
6. $\text{S}_8(\text{s}) + \text{NO}_3^{-}(\text{aq}) \longrightarrow \text{SO}_3^{2-}(\text{aq}) + \text{NO}(\text{g})$
7. $\text{H}_3\text{AsO}_4(\text{aq}) + \text{Zn}(\text{s}) \longrightarrow \text{AsH}_3(\text{g}) + \text{Zn}^{2+}(\text{aq})$
8. $\text{P}_4(\text{s}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) \longrightarrow \text{H}_3\text{PO}_4(\text{aq}) + \text{Cr}^{3+}(\text{aq})$

Redox Reactions in Basic Solution

1. $\text{Al}(\text{s}) + \text{MnO}_4^{-}(\text{aq}) \longrightarrow \text{MnO}_2(\text{s}) + \text{Al}(\text{OH})_4^{-}(\text{aq})$
2. $\text{NO}_2^{-}(\text{aq}) + \text{Al}(\text{s}) \longrightarrow \text{NH}_3(\text{aq}) + \text{AlO}_2^{-}(\text{aq})$
3. $\text{Cr}(\text{s}) + \text{CrO}_4^{2-}(\text{aq}) \longrightarrow \text{Cr}(\text{OH})_3(\text{s})$
Note: $\text{Cr}(\text{OH})_3$ is found in BOTH half reactions!
4. $\text{MnO}_4^{-}(\text{aq}) + \text{S}^{2-}(\text{aq}) \longrightarrow \text{MnO}_2(\text{s}) + \text{SO}_3^{2-}(\text{aq})$
5. $\text{Cl}_2(\text{aq}) + \text{Br}_2(\text{l}) \longrightarrow \text{OBr}^{-}(\text{aq}) + \text{Cl}^{-}(\text{aq})$
6. $\text{H}_2\text{O}_2(\text{aq}) + \Gamma^{-}(\text{aq}) \longrightarrow \text{IO}_3^{-}(\text{aq})$
Note: IO_3^{-} is found in both half reactons!
7. $\text{NO}_3^{-}(\text{aq}) + \text{NH}_3(\text{aq}) \longrightarrow \text{NO}_2^{-}(\text{aq})$
8. $\text{S}_8(\text{aq}) + \text{MnO}_4^{-}(\text{aq}) \longrightarrow \text{SO}_4^{2-}(\text{aq}) + \text{MnO}_2(\text{s})$