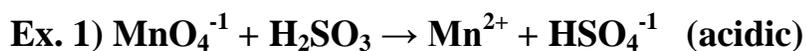


## Step by Step: Oxidation Numbers and Balancing Redox reactions.



Rules for Oxidation Numbers (for individual atoms):

1) Any free element (without a charge) is zero. Ex.)  $\text{Zn}_{(s)}$ ,  $\text{N}_{2(g)}$

2) Monoatomic ions have their charge. Ex.)  $\text{Zn}^{2+}_{(aq)}$ ,  $\text{F}^{-1}_{(aq)}$

3) Hydrogen is always +1, unless it is alone with a positive metal, then the H is -1.

Ex.)  $\text{NaH} = \text{Na}^{+1} \text{H}^{-1}$

4) Oxygen is always -2, unless it is in a peroxide, then O is -1.

Ex.)  $\text{H}_2\text{O}_2 = \text{H}^{+1} \text{O}^{-1}_2$

5) Fluorine is always -1.

6) The sum of oxidation numbers in a group of elements equals the charge of the group.

Ex.)  $\text{NaCl}$  so  $\text{NaCl}^{-0}$ ,  $\text{SO}_4^{2-}$  so  $\text{SO}_4^{-2}$

\*\* Remember charges on Periodic Table Columns, for single element ions.

(1<sup>st</sup> column = +1, 2<sup>nd</sup> column = +2 etc.)

**$\text{MnO}_4^{-1}$  : Oxygen is always -2 (Rule 4). If there are 4 oxygens, then  $4 \times (-2) = -8$  for all of the oxygens.**

**$\text{Mn} + (-8 \text{ for the oxygens}) = -1$  (It must all add up to -1, since that is the charge given for the ion.)**

**Mn must equal +7.**

**$(\text{Mn}^{+7} \text{O}^{-2}_4)^{-1}$**

**+7 -8 = -1 It works!**

**$\text{H}_2\text{SO}_3$  : Hydrogen is always +1 (Rule 3). Oxygen is always -2 (Rule 4). If there are 2 hydrogens  $2 \times (+1) = +2$ . If there are 3 oxygens,  $3 \times (-2) = -6$ .**

**$(+2 \text{ for the hydrogens}) + \text{S} + (-6 \text{ for the oxygens}) = 0$  (This compound is neutral, no charge written above it.)**

**S must equal +4.**

**$\text{H}^{+1}_2 \text{S}^{+4} \text{O}^{-2}_3$**

**+2 +4 -6 = 0 It works!**

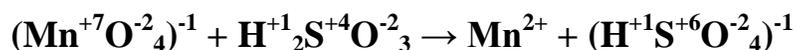
**$\text{Mn}^{2+}$  is alone and has a charge, so it must add up to +2 (Rule 2).**

$\text{HSO}_4^{-1}$  : Hydrogen is always +1 (Rule 3). Oxygen is always -2 (Rule 4). If there are 4 oxygens,  $4X(-2) = -8$ .

$(+1 \text{ for the hydrogen}) + \text{S} + (-8 \text{ for the oxygens}) = -1$  (This compound is neutral, no charge written above it.)

S must equal +6.

$(\text{H}^{+1}\text{S}^{+6}\text{O}^{-2}_4)^{-1}$   
 $+1 +6 -8 = -1$  It works!



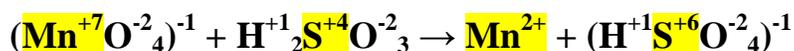
**\*\*For more practice on oxidation numbers go to the first part of each of the following problems!**

### Steps for balancing redox:

- 1) Write the half reactions. (one for oxidation, one for reduction)
- 2) For each half reaction:
  - a) Balance the electrons, by looking at the oxidation number of the elements that are oxidizing or reducing. (Make sure these elements are first balanced!)
  - b) Balance all other atoms except H and O. ( If necessary, put in other compounds from the original reaction.)
  - c) Balance O by adding  $\text{H}_2\text{O}$ .
  - d) Balance H by adding  $\text{H}^+$ .
  - e) Make sure all elements and charges balance.
- 3) Multiply the half reactions by numbers so that each has the same number of electrons.
- 4) Add half reactions and cancel what you can.
- 5) **If in basic solution:**
  - a) Notice how many  $\text{H}^+$  are left. Add the same amount of  $\text{OH}^-$  as  $\text{H}^+$ , to both sides of the reaction.
  - b) The  $\text{OH}^-$  and  $\text{H}^+$  that are on the same side make  $\text{H}_2\text{O}$ .
  - c) Cancel any  $\text{H}_2\text{O}$  that will cancel.
- 6) Make sure all charges and atoms balance.

We must balance these in this strange method to get the charges to balance. If we balanced these using the old method, we could get the elements to balance, but the charges would not balance. We have to have the charges balance, since the oxidizing elements are giving electrons to the reducing elements. These numbers of electrons will eventually have to be equal (Step #3). If the two half reactions are physically separated and a wires placed between them, a battery can be made. These are battery reactions, since they transfer electrons and can give off current.

**Look at the charges on the single elements! Not the total charges of all hydrogens or oxygens, only the charge on a single element of that type!**



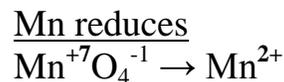
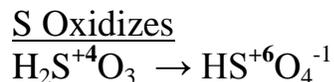
Find the elements that have changing oxidation numbers going from the left side of the reaction to the right side of the reaction!

Sulfur goes from +4 to +6, so it is **oxidizing**; it is getting more **positive**.

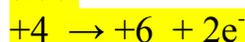
Mn goes from +7 to +2, so it is **reducing**; it is getting **smaller (reducing)**.

Write a reaction using only the sulfur compounds (for oxidizing) and write a reaction using only the manganese compounds (for reducing).

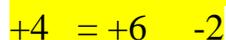
### Step 1:



Looking at only the +4 and +6 on the sulfurs, add electrons to balance the charge. (Make the charge the same on each side of the reaction.) Always add the electrons to the more positive side.



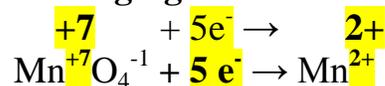
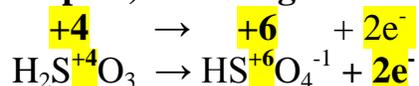
Now +4 on the left will equal the +6 + (-2 charge from the electrons) on the right.



Do the same for the Mn: only look at the +7 and the +2 on the Mn. Add electrons to the more positive side.



### Step 2a): Add negative electrons to balance changing oxidation #'s.



### Step 2b): all balanced except H, O

There are no other elements that need balanced other than Oxygen and Hydrogen, so step 2b can be skipped.

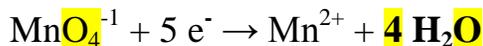
Next we will balance Oxygen by adding water (H<sub>2</sub>O). This will mess up the Hydrogens, but we will fix them in the next step.

We want to add water to whichever side needs Oxygen!

**For Oxidizing:** Looking below we have 3 oxygens on the left side (H<sub>2</sub>SO<sub>3</sub>) and 4 oxygens on the right side (HSO<sub>4</sub><sup>-1</sup>). So we need one more oxygen on the left side to make 4 total oxygens on each side. So we add one H<sub>2</sub>O to the left side.

**For Reducing:** Looking below we have 4 oxygens on the left side (MnO<sub>4</sub><sup>-1</sup>) and none on the right side. So we need 4 oxygens on the right side, add 4 H<sub>2</sub>O to the right side.

**Step 2c): Add H<sub>2</sub>O to balance O.**



Next we balance the Hydrogens by adding H<sup>+</sup> to whichever side needs more hydrogens.

**For Oxidizing:** There are 4 hydrogens on the left (H<sub>2</sub>O + H<sub>2</sub>SO<sub>3</sub>) and 1 hydrogen on the right (HSO<sub>4</sub><sup>-1</sup>). So we need to add 3 hydrogens to the right side to have a total of 4 hydrogens on each side.

**For Reducing:** There are no hydrogens on the left and 8 hydrogens on the right. We need to add 8 hydrogens on the left side.

**Step 2d): Add H<sup>+</sup> to balance H.**

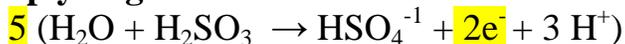


Next we need to have the same number of electrons on each side, so that we can cancel them in the next step!

**For Oxidizing:** There are 2 electrons.

**For Reducing:** There are 5 electrons. We need to make 10 total (lowest common number). So the whole **oxidizing** reaction is multiplied by 5. (5 times each compound to keep it balanced.) The whole **reducing** reaction is multiplied by 2.

**Step 3: Multiply to get common # of electrons.**



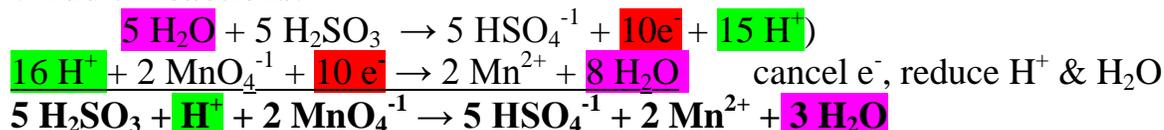
Next we cancel or reduce whatever we can to make the reaction simpler and then add it up. We can cancel or reduce, if they are on opposite sides of the arrows. If they are on the same side of the arrows, they will add together (see example 3).

The 10 electrons completely cancel from each side (subtract 10 from each side).

15 H<sup>+</sup> will subtract from each side, leaving only 1H<sup>+</sup> on the left side.

5 H<sub>2</sub>O will subtract from each side, leaving only 3 H<sub>2</sub>O on the right side.

**Step 4: Add 1/2 reactions.**



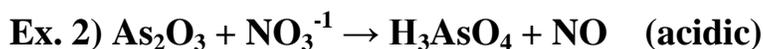
The order does not matter as long as they are on the correct side of the arrow.

The elements and charges should all be equal on both sides.

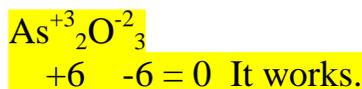
See below for the check on the charges:



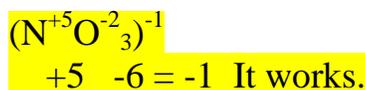
It balances chargewise! **You do not look at the individual charges you used to get the electrons, only at the charges on whole ions.** Make sure you did not lose any charges from the original problem!



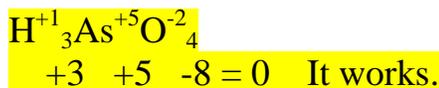
**$\text{As}_2\text{O}_3$ :** Oxygen is -2 (Rule 4). 3 oxygens times -2 equals -6. So the As side must equal +6 (since the compound is neutral), but there are 2 As, so each will equal +3.



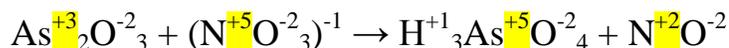
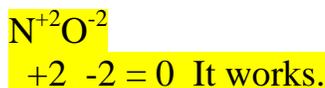
**$\text{NO}_3^{-1}$ :** Oxygen is -2 (Rule 4). 3 oxygens times -2 equals -6. But the whole nitrate ion has to add up to -1. So nitrogen will be +5.



**$\text{H}_3\text{AsO}_4$ :** Hydrogen is +1 (Rule 3). Oxygen is -2 (Rule 4). 3 hydrogens equals +3 and 4 oxygens equals -8. For the compound to be neutral As must be +5.

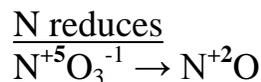
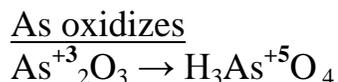


**NO:** Oxygen is -2 (Rule 4). 1 oxygen times -2 equals -2, so nitrogen must be +2 to make a neutral compound.



As is oxidizing, since it goes from +3 to +5 (getting more positive). (Remember you are looking at the single charges for each element, not the total/overall charge.)

N is reducing, since it goes from +5 to +2 (getting smaller). Now write the half reactions using the compounds containing those elements.



**For Oxidizing:** \*\*This time we need to balance the arsenic (As) before adding the electrons, since there are 2 As on the left side and only 1 As on the right side.

When we add the 2 in front of the compound containing As, the charge will also be affected.

On the left side of the reaction we have 2 As that have a +3 charge equaling +6.

On the right side of the reaction we have 2 As that have a +5 charge equaling +10.

We will need to add 4 negative electrons to the most positive side, the right side. Now we have a charge of +6 on both sides of the reaction.

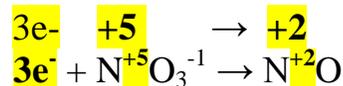
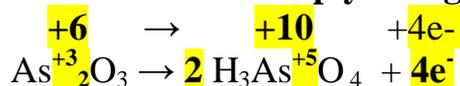


**For Reducing:** The nitrogens are balanced, so we can go straight to adding electrons. The nitrogen on the left has a +5 charge and the nitrogen on the right has a +2 charge, so we need to add 3 negative electrons to the more positive +5 side. Now we have a charge of +2 on both sides of the reaction.



**\*\* Balance As before doing electrons,**

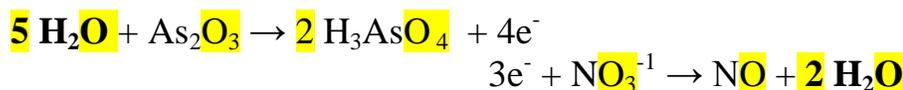
**so then multiply charges by 2, since there are 2 Arsenics!!**



Next we add H<sub>2</sub>O to balance Oxygen.

**For oxidizing:** we have 3 oxygens on the left and 8 oxygens (2 X 4 O) on the right. We need 5 more oxygens on the left side.

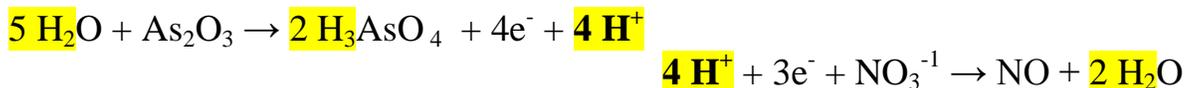
**For reducing:** we have 3 oxygens on the left side and 1 oxygen on the right, so we need to add 2 more oxygens to the right side.



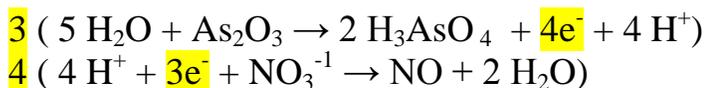
Next we add H<sup>+</sup> to balance the hydrogens:

**For oxidizing:** we have 10 hydrogens on the left and on the right we have 6, so we need 4 more on the right.

**For reducing:** we have 0 hydrogens on the left and on the right we have 4, so we need 4 on the right.



Next we need to cancel the electrons: we have 4 electrons for oxidizing and 3 for reducing. We can make 12, by multiplying by 3 and by 4.

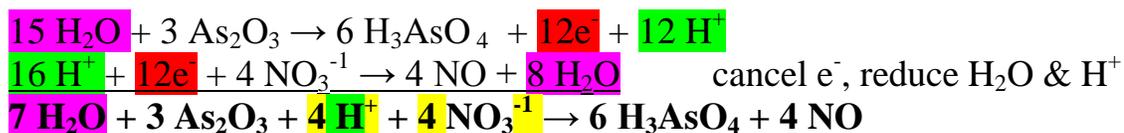


Then we need to cancel and reduce:

The electrons cancel.

We can cancel (subtract) 8 H<sub>2</sub>O from each side.

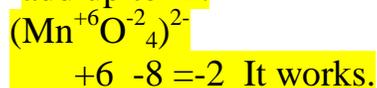
We can cancel (subtract) 12 H<sup>+</sup> from each side.



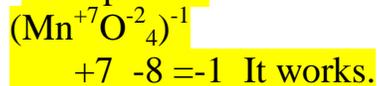
The charge balances: +4 -4 = 0 zero charge



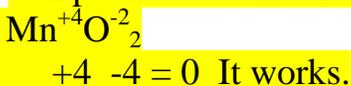
**MnO<sub>4</sub><sup>2-</sup>:** Oxygen is -2 (Rule 4). 4 oxygens times -2 equals -8. Mn must be +6 for the ion to add up to -2.



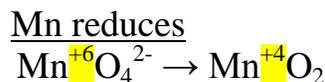
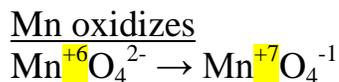
**MnO<sub>4</sub><sup>-1</sup>:** Oxygen is -2 (Rule 4). 4 oxygens times -2 equals -8. Mn must be +7 for the ion to add up to -1.



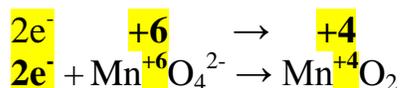
**MnO<sub>2</sub>:** Oxygen is -2 (Rule 4). 2 oxygens times -2 equals -4. Mn must be +4 for the compound to be neutral.



Mn is oxidizing (+6 to +7) and Mn is reducing (+6 to +4).



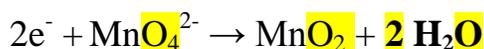
Mn is already balanced in each reaction. (There is one on each side.) So we can go straight to balancing the electrons. (Remember electrons are added to the most positive side, so that we have the same charge on both sides.)



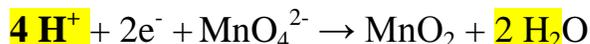
**For oxidizing:** this 1/2 reaction is already balanced and needs no oxygens or hydrogens.



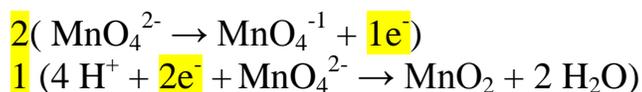
**For reducing:** we need to balance oxygens first by adding H<sub>2</sub>O. We need 2 waters on the right so that we have a total of 4 oxygens on each side.



**For reducing:** we need to balance hydrogen by adding H<sup>+</sup>. We need 4 H<sup>+</sup> on the left, so that we have a total of 4 hydrogens on each side.



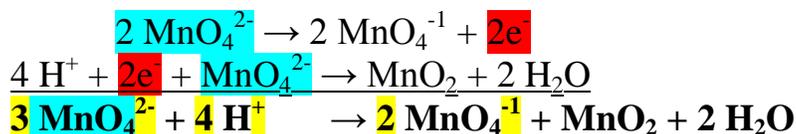
Next we need to cancel the electrons. Oxidizing needs multiplied by 2, so that we have 2 electrons in each reaction.



Then we can cancel electrons.

The MnO<sub>4</sub>'s do not cancel across (left to right), since the charges are different.

The MnO<sub>4</sub><sup>2-</sup> that are on the same side add together!



Charge balances:  $-6 + 4 = -2 \quad -2$

Basic reactions are the same as acid, except they have one extra step at the end!



$\text{NO}_2^{-1}$ : Oxygen is -2 (Rule 4). 2 oxygens times -2 equals -4, but we need the ion to equal -1, so nitrogen will be +3.



+3 -4 = -1 It works.

**Al**: aluminum is alone and there is no charge written, so it is zero (Rule 1).

$\text{NH}_3$ : Hydrogen is +1 (Rule 3). 3 hydrogens times +1 equals +3, but we need a neutral compound, so nitrogen will be -3.



-3 +3 = 0 It works.

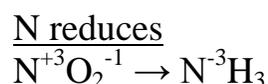
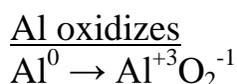
$\text{AlO}_2^{-1}$ : Oxygen is -2 (Rule 4). 2 oxygens times -2 equals -4, but we need the ion to equal -1, so aluminum will be +3.



+3 -4 = -1

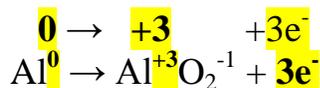


Al oxidizes (0 to +3), N reduces (+3 to -3).

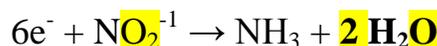
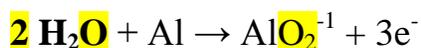


Al and N are balanced. There is one on each side, so we can do the electrons. Add the electrons to the more positive side, so that we have equal charges on both sides.

\*\*Note on the nitrogens we go from +3 to -3, think of a number line +3, +2, +1, 0, -1, -2, -3 we need to go 6 positions, so we need 6 electrons (-6 + 3 = -3).

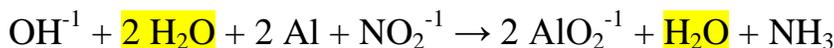


Next we add  $\text{H}_2\text{O}$  to balance the oxygens.

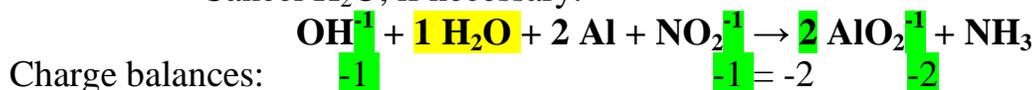


Next we add  $\text{H}^+$  to balance the hydrogens.





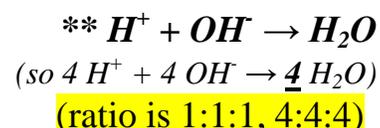
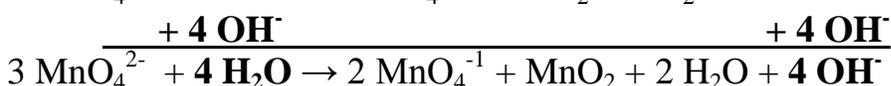
Cancel H<sub>2</sub>O, if necessary.



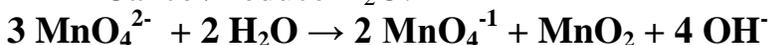
Ex. 5) Do Ex. 3) as if it were BASIC.



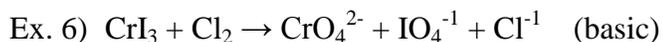
Note the number of H<sup>+</sup>, and add this amount of OH<sup>-</sup> to both sides.



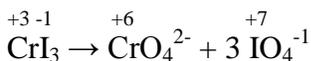
Cancel/Reduce H<sub>2</sub>O:



## Redox Reactions with more than one element oxidizing or reducing:



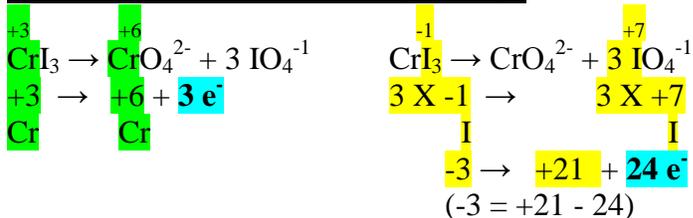
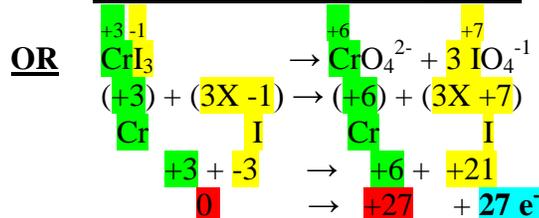
Oxidation



Reduction

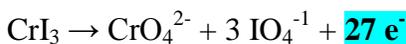
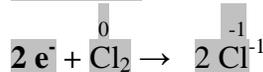


For oxidation, the number of electrons can be calculated separately for the Cr and I reactions (then totaling the electrons for both reactions) **or** the Cr and I reactions can be combined together so that the electrons are totaled together in the combined reaction.

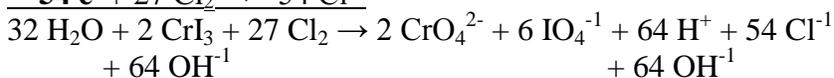
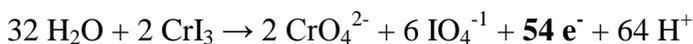
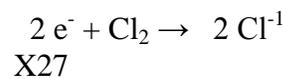
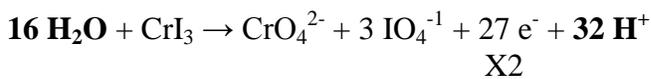
**Oxidation Reactions Done Separately:****Oxidation Combined Together:**

Combine the  $3 \text{e}^-$  and  $24 \text{e}^-$  electrons, so  $27 \text{e}^-$  on the right.

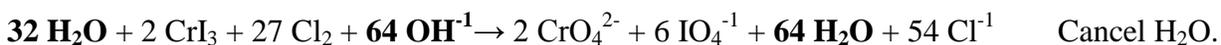
Above we also get  $27 \text{e}^-$ .

OxidationReduction

Then do your normal balancing with  $\text{H}_2\text{O}$  and  $\text{H}^+$ .



Basic, so add  $\text{OH}^-$ .



Check charge:

$$-64 \quad =? \quad -4 \quad -6 \quad -54 \quad \checkmark$$

Check elements:

$$2 \text{Cr}, 6 \text{I}, 54 \text{Cl}, 64 \text{O}, 64 \text{H} \quad =? \quad 2 \text{Cr}, 8+24+32 = 64 \text{O}, 6 \text{I}, 54 \text{Cl}, 64 \text{H} \quad \checkmark$$

**\*End of Notes\***