

Since O has oxidation number = -2, $(7)(-2) + 2(\text{Cr}) = -2$

$$7(-2) + 2(\text{Cr}) = -2$$

$$2(\text{Cr}) = 12$$

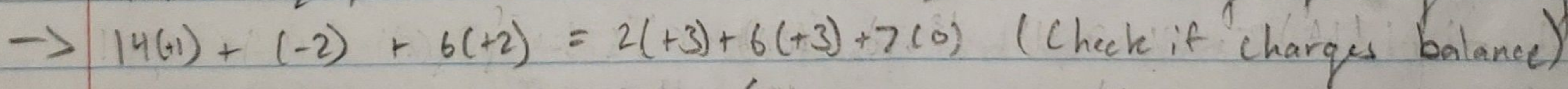
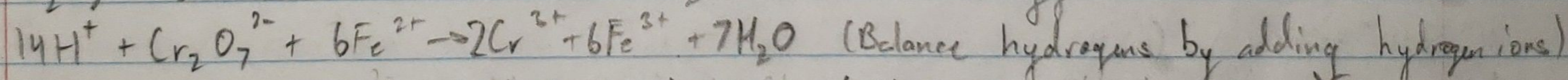
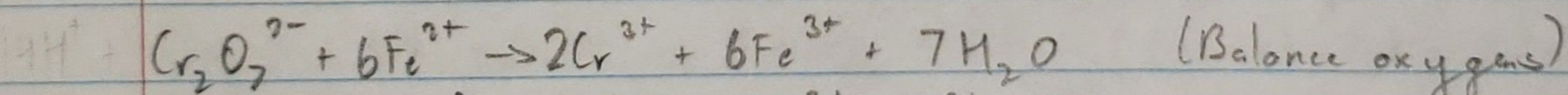
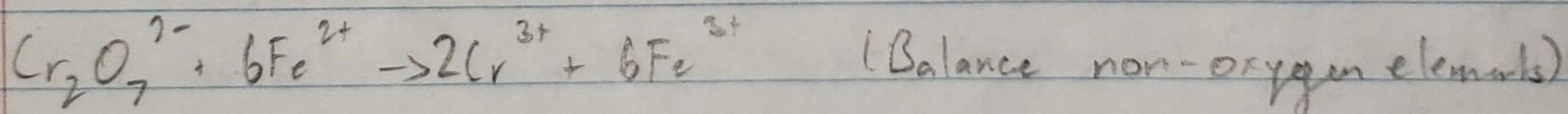
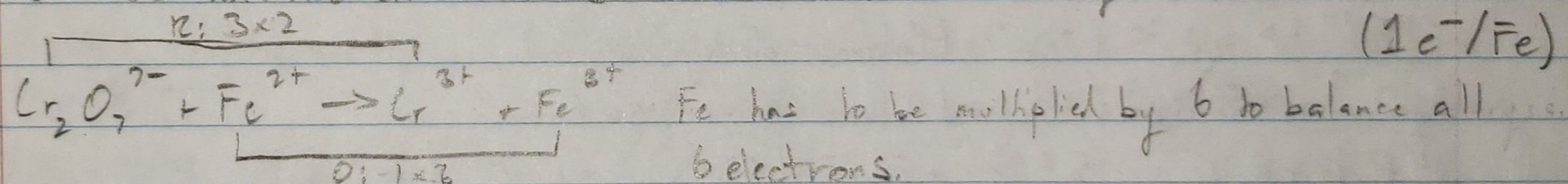
$\text{Cr} = 6$, \therefore Cr has oxidation number +6

Cr^{3+} means that Cr as a product, has an oxidation number of +3. Cr is reduced since it gains 6 electrons. (There are 2 Cr in the reactants)

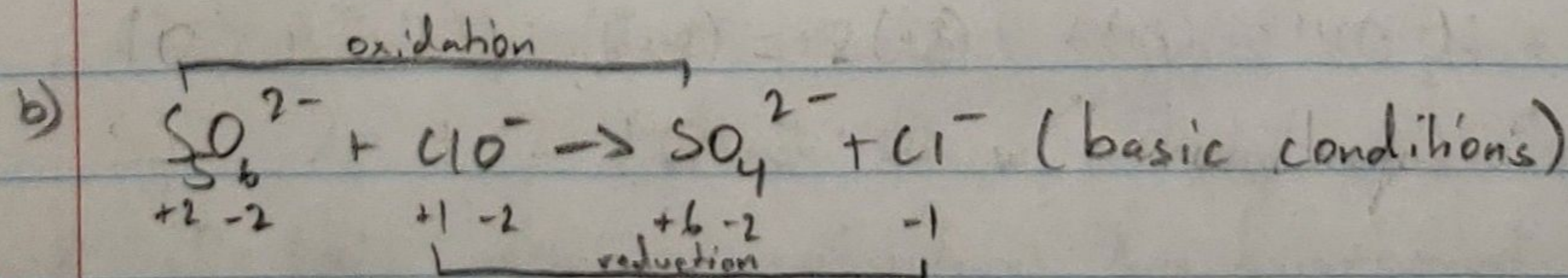
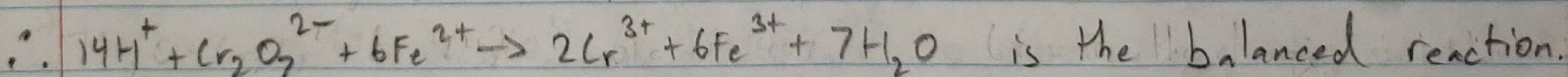
$(3e^-/\text{Cr})$, $(6e^-/2\text{Cr})$

Fe loses an electron when its oxidation number changes from +2 to +3 (oxidized)

$(1e^-/\text{Fe})$



$24 = 24 \checkmark$



For SO_3^{2-} :

O has oxidation number = -2, S has oxidation number: $6(-2) + 5(\text{S}) = -2$

$$-12 + 5(\text{S}) = -2$$

For SO_4^{2-} :

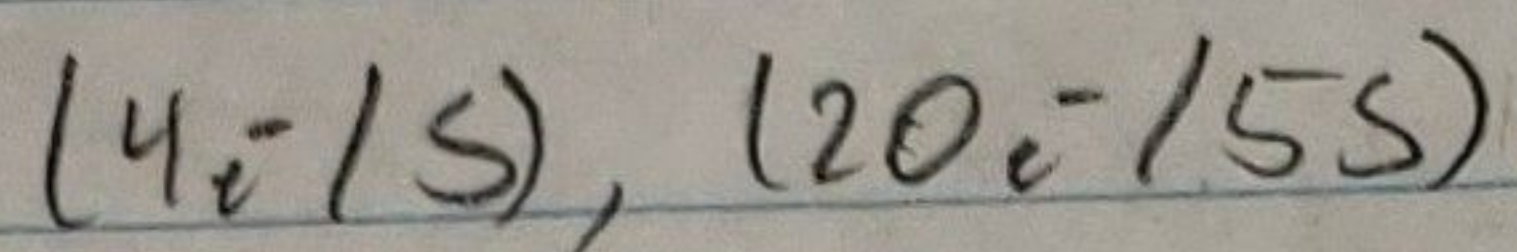
$\text{S} = +2$

O has oxidation number = -2, S has oxidation number: $4(-2) + \text{S} = -2$

$\text{S} = +6$

1b) con't

Each S has a change in oxidation number of +2 to +6. Each S loses 4 electrons
for a total of 20 electrons. (Oxidation)



For ClO^- :

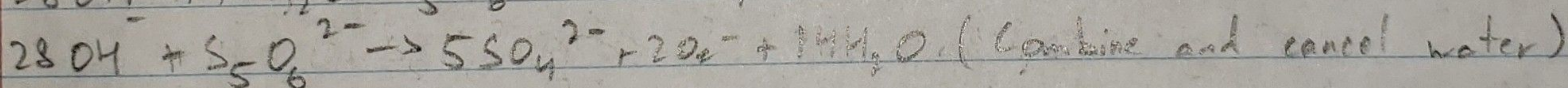
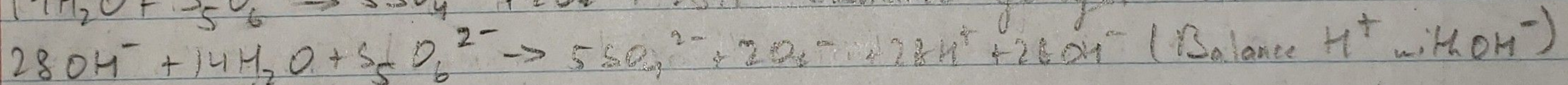
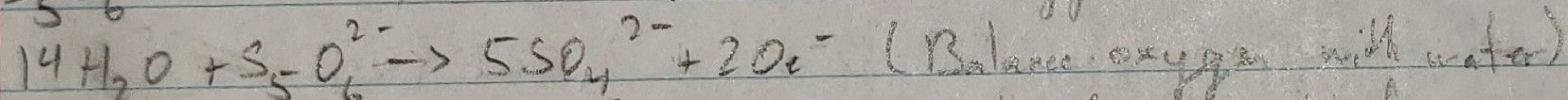
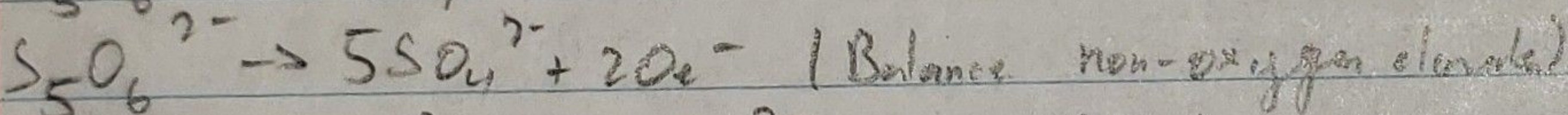
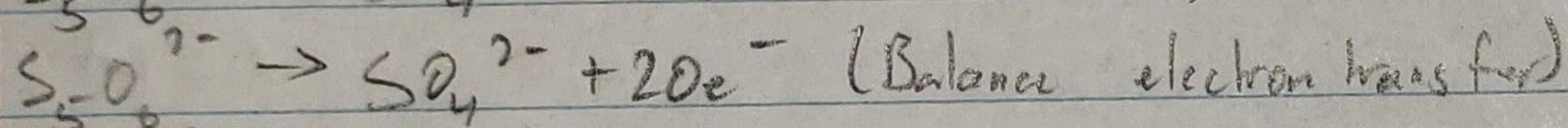
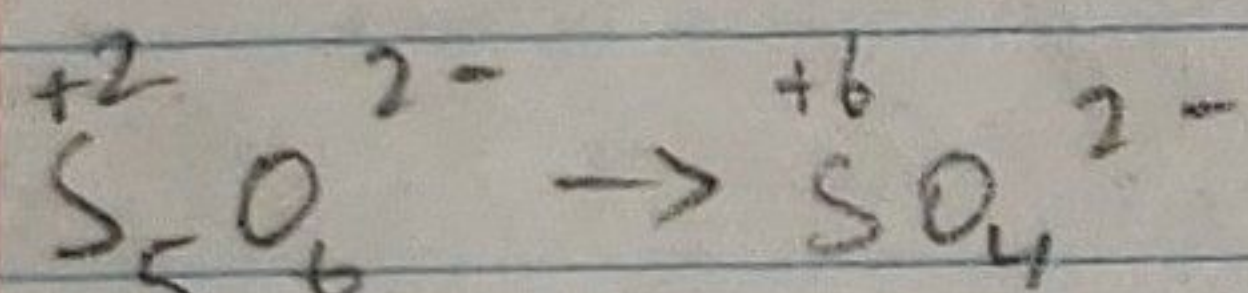
O has an oxidation number of -2, Cl has an oxidation number: $-2 + \text{Cl} = -1$
 $\text{Cl} = +1$.

For Cl^- :

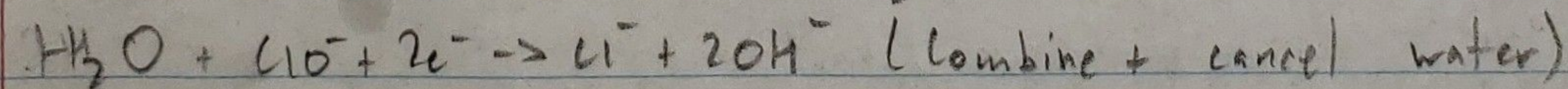
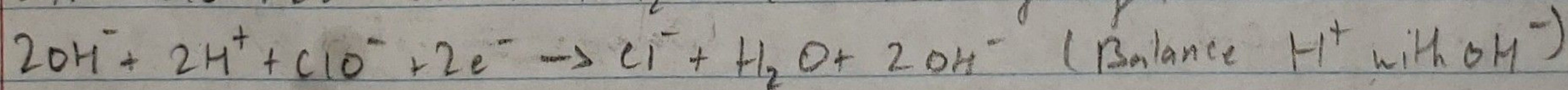
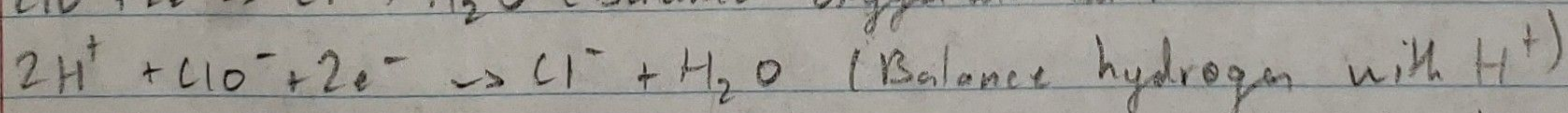
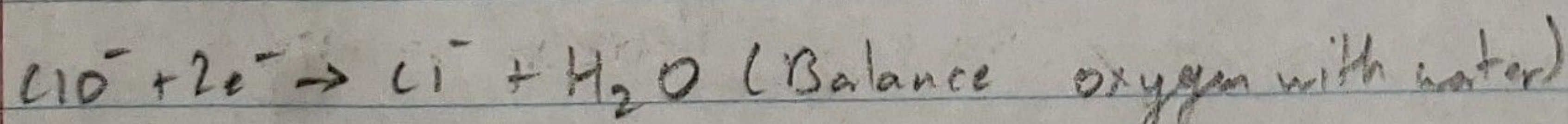
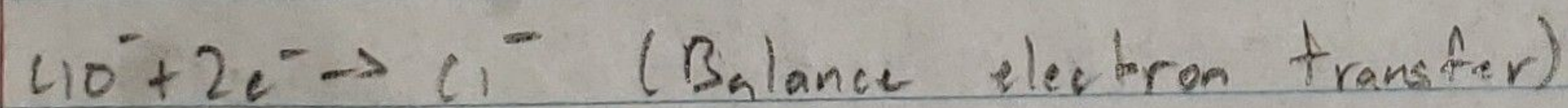
Cl has an oxidation number of -1.

Each Cl has a change in oxidation number from +1 to -1. Cl gains 2 electrons (reduced)
 $(2e^-/\text{Cl})$

O:



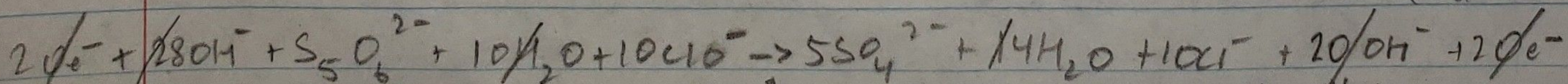
Cl: $\text{ClO}^- \rightarrow \text{Cl}^-$



\therefore oxidation reaction has $20e^-$ transferred and reduction reaction has $2e^-$ transferred,

multiply reduction reaction by 10 and add half-reactions:

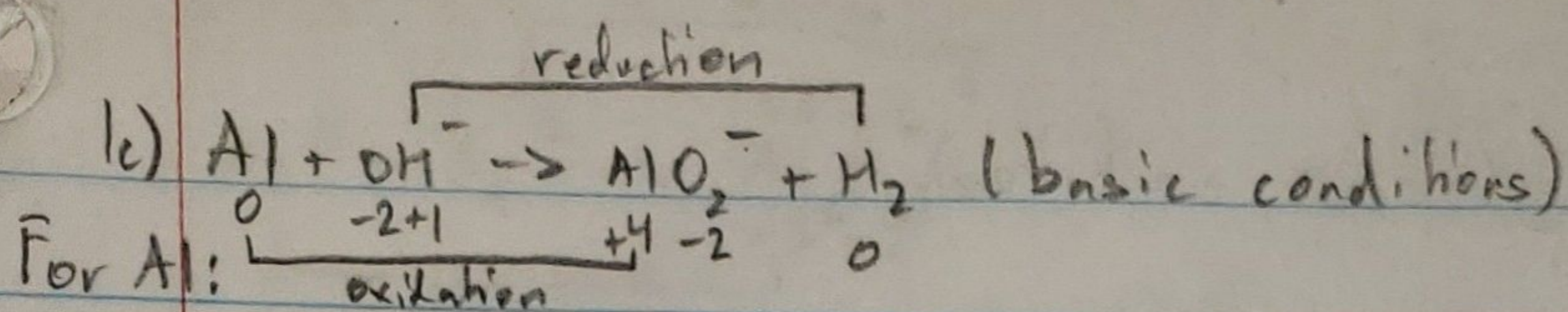
(Cancel like terms)



$\therefore 8\text{OH}^- + \text{S}_5\text{O}_6^{2-} + 10\text{ClO}^- \rightarrow 5\text{SO}_4^{2-} + 4\text{H}_2\text{O} + 10\text{Cl}^-$ is the balanced reaction
in basic conditions

$$8(-1) + (-2) + 10(-1) = 5(-2) + 4(0) + 10(-1)$$

$$-20 = -20 \checkmark \text{ Charges balance}$$



Al has an oxidation number of 0

Al in AlO_2^- has an oxidation number: $\text{Al} + 2(-2) = -1$ (oxygen has an oxidation number of -2)
 $\text{Al} = +3$

Al loses 3 electrons (oxidation) ($3e^-/\text{Al}$)

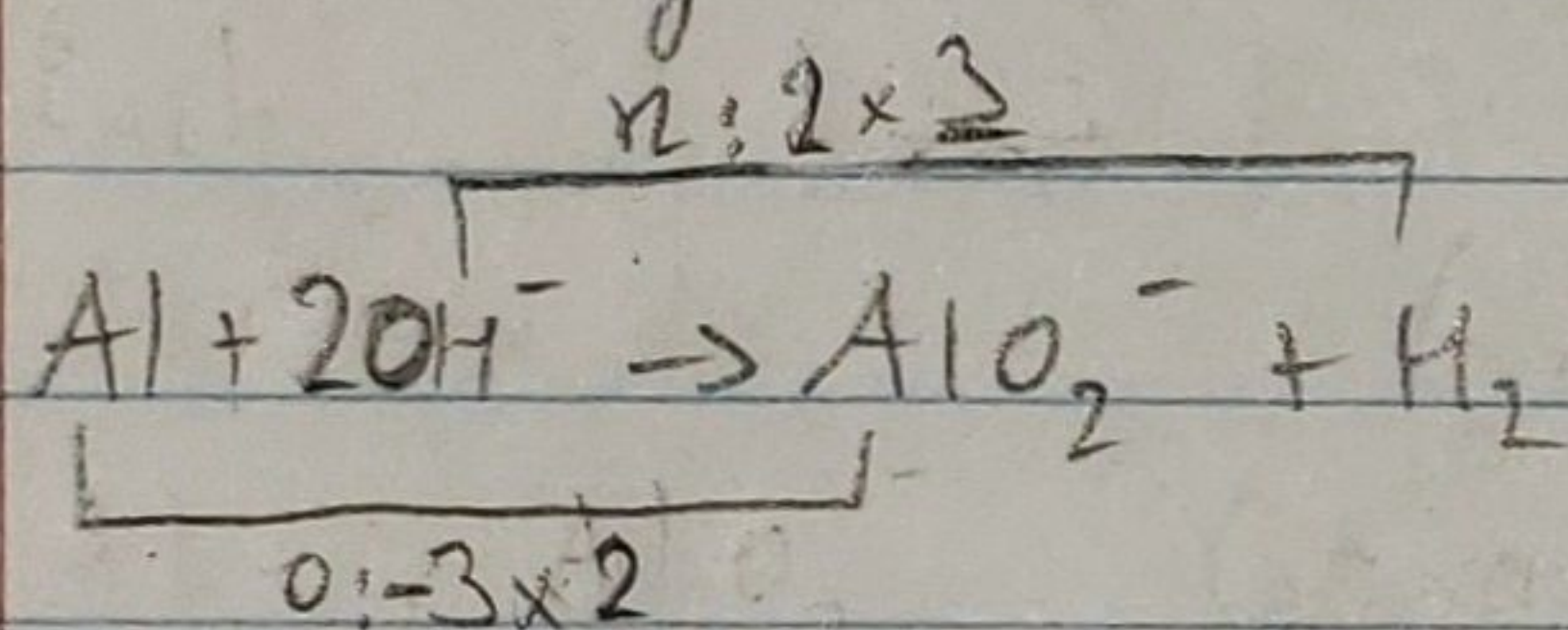
For H:

H has an oxidation number: $\text{H} + (-2) = -1$

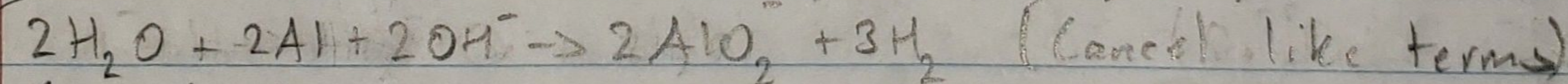
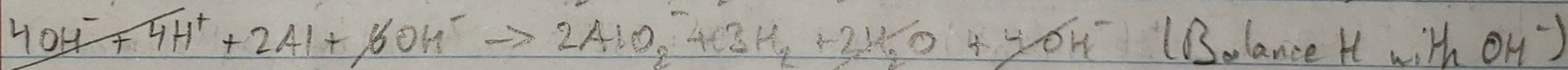
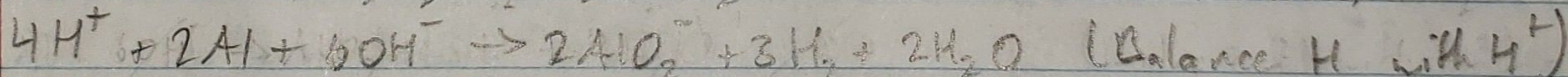
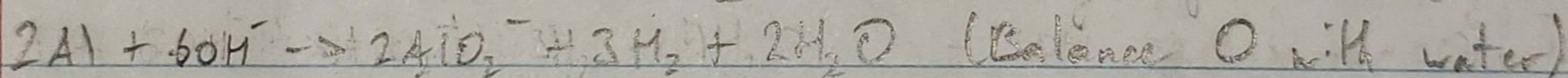
$\text{H} = +1$ in OH^- .

H has an oxidation number = 0 in H_2 .

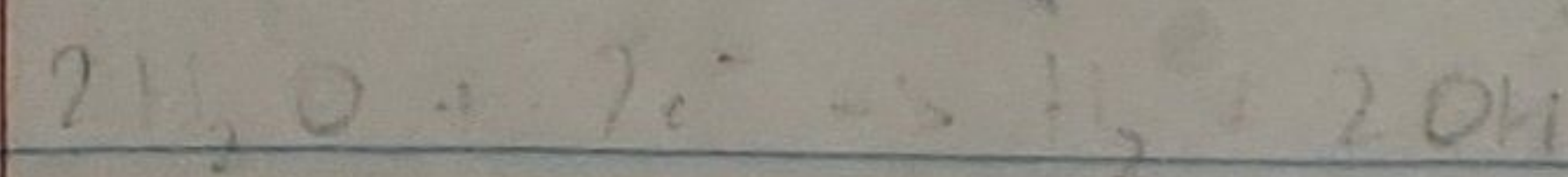
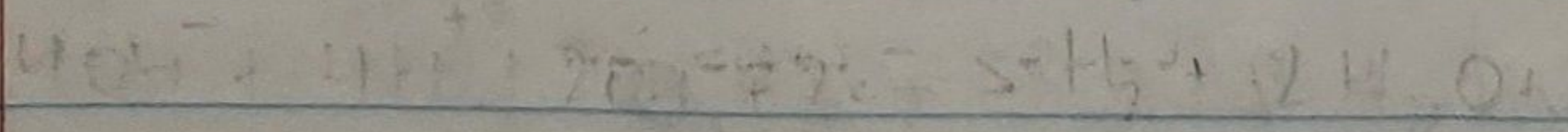
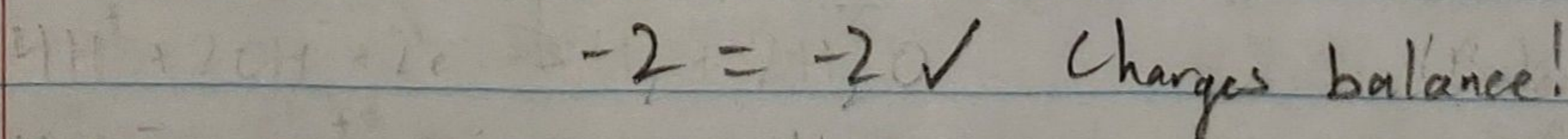
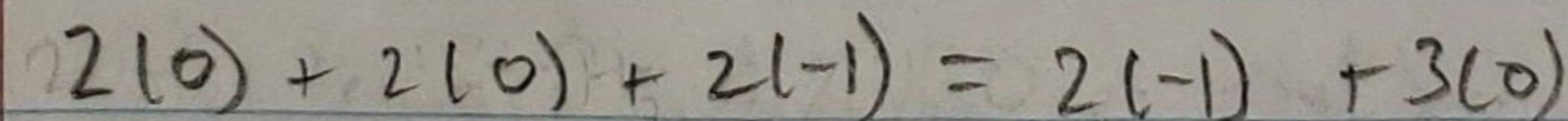
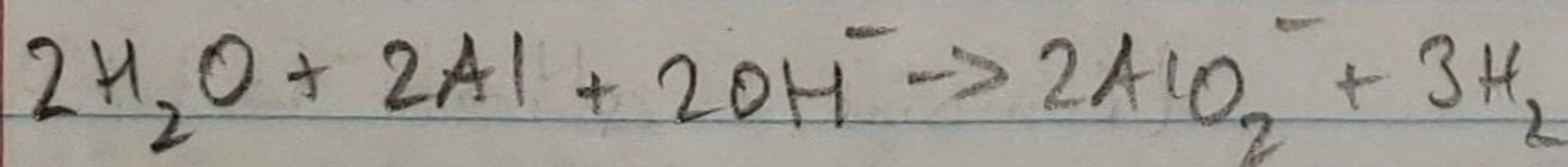
Each H gains 1 electron (reduction) for a total of 2 electrons. ($1e^-/\text{H}$), ($2e^-/2\text{H}$)



Coefficients in front of Al and AlO_2^- are multiplied by 2 and coefficients in front of OH^- and H_2 are multiplied by 3 since $\text{LCM}(3,3) = 6$.

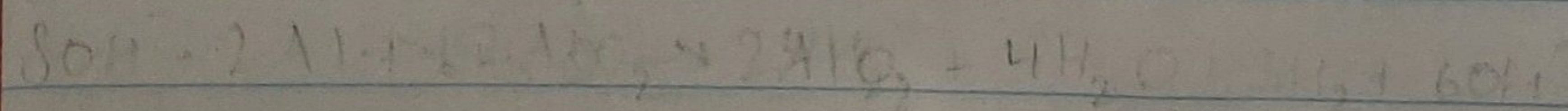


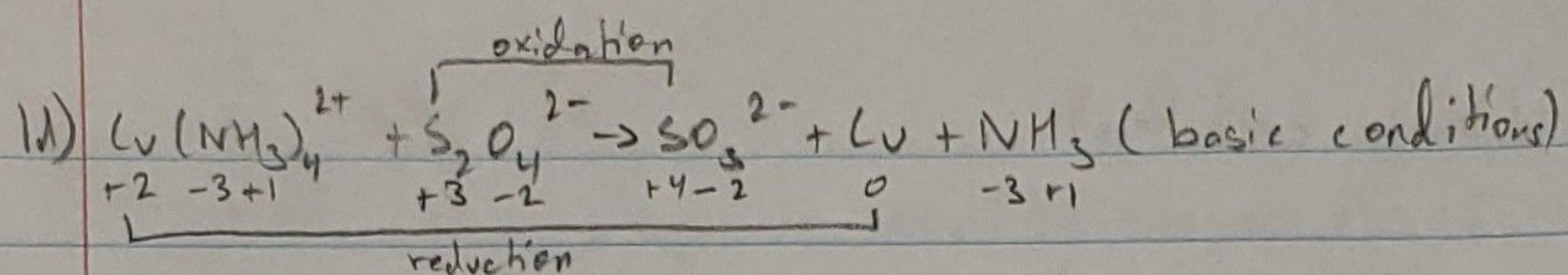
\therefore The balanced reaction in basic conditions is:



Multiply reduction and oxidation reactions by 2 to balance electrons.

All half-reactions:





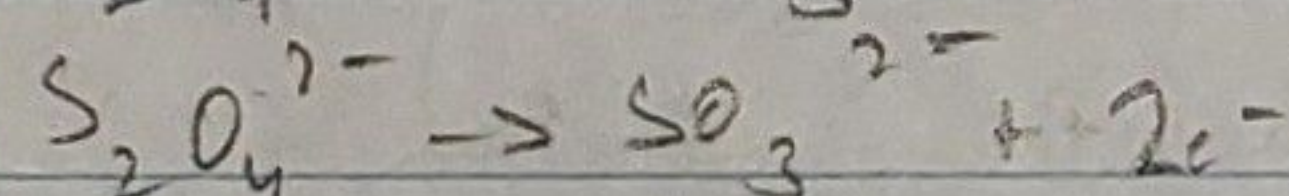
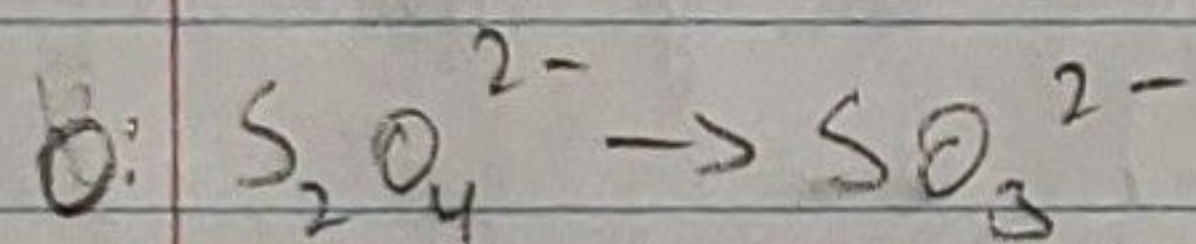
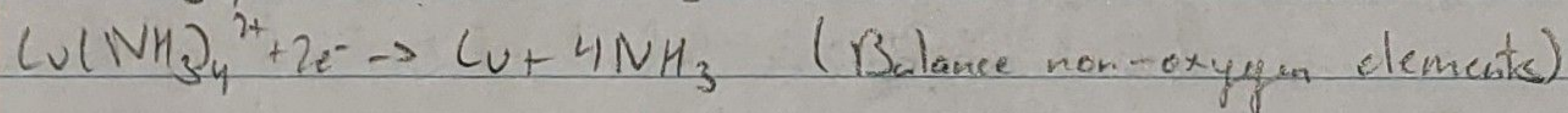
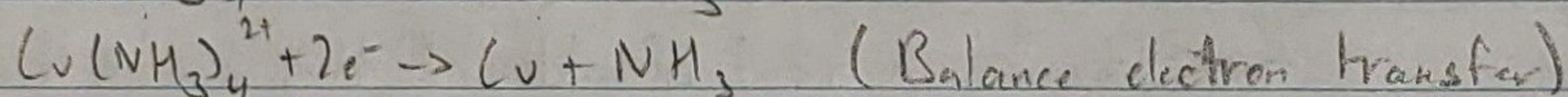
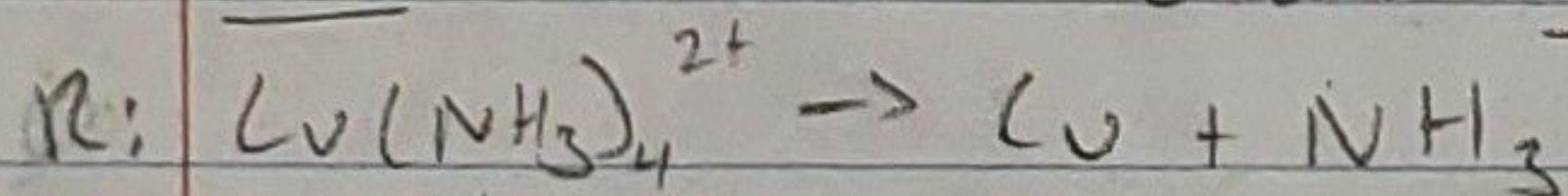
For Cu:

Cu has a +2 oxidation number in $\text{Cu(NH}_3)_4^{2+}$ (NH_3 has a total oxidation number of 0)
 Cu has an oxidation number = 0 in Cu.
 Cu gains 2 electrons (reduction) ($2e^-/\text{Cu}$)

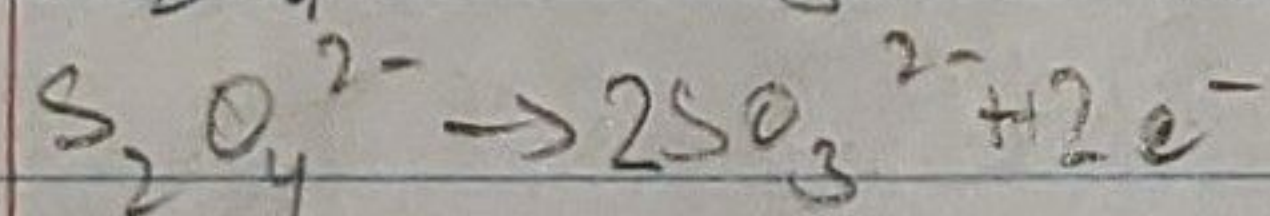
For S: S has oxidation number: $2(S) + 4(-2) = -2$ in $\text{S}_2\text{O}_4^{2-}$
 $S = +3$

S has oxidation number: $S + 3(-2) = -2$ in SO_3^{2-}
 $S = +4$

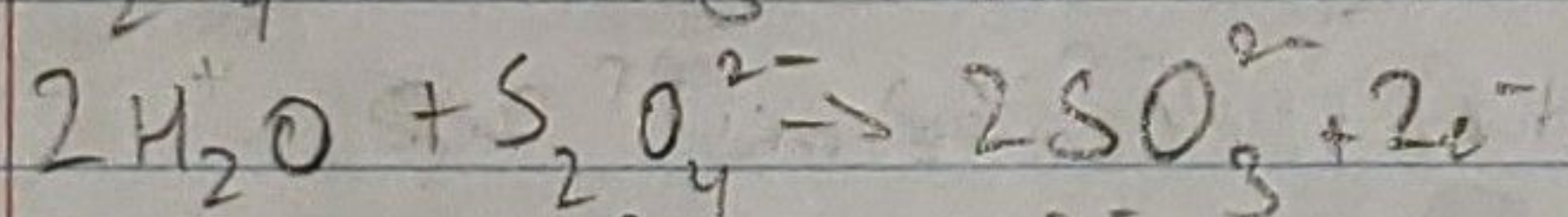
Each S loses 1 electron (oxidation) ($1e^-/\text{S}$), Total: $2e^-/2\text{S}$



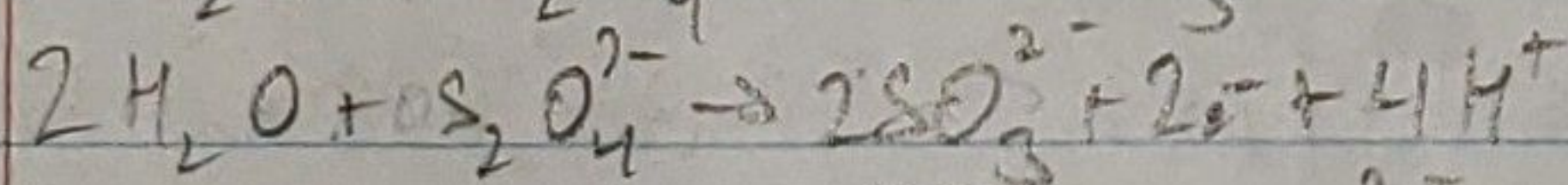
(Balance electron transfer)



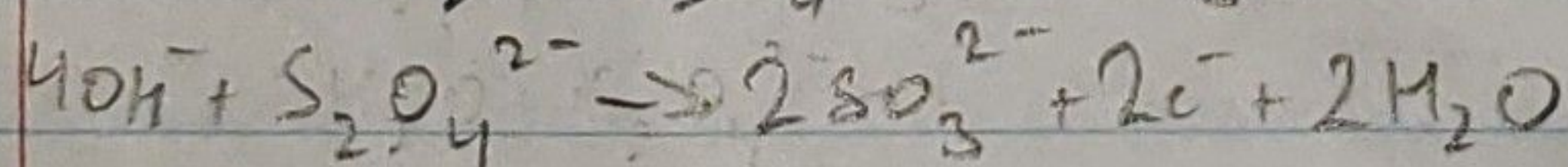
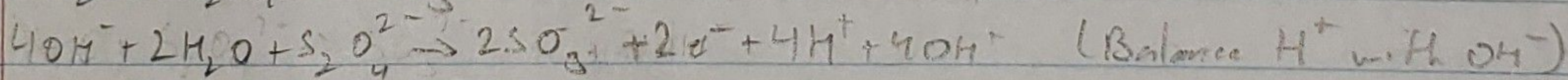
(Balance non-oxygen elements)



(Balance O with water)

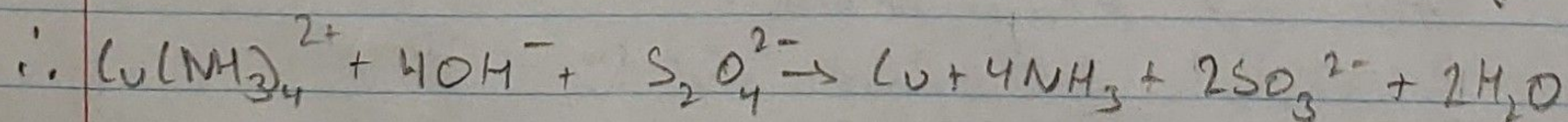


(Balance H with H^+)



(Combine + cancel water)

Add half-reactions since electron gain/loss is already equal.



$$+2 + 4(-1) + (-2) = 0 + 4(0) + 2(-2) + 2(0)$$

$$-4 = -4 \quad \checkmark \text{ Charges balance!}$$