XV) Applied Electrochemistry

Breathalyzer

Police use a breathalyzer to test for alcohol consumption. The breathalyzer consists of a redox reaction that includes a colour change. The intensity of green colour produced is measured to find the amount of alcohol in the blood. The colour change is due to the orange dichromate ion ($Cr_2O_7^{2-}$) reacting to produce the green chromium ion (Cr^{3+}) due to a reaction with ethanol, C_2H_5OH, the alcohol used in beverages.

discussed earlier in redox titrations

$$C_2H_5OH + K_2Cr_2O_7 + H_2SO_4 \Rightarrow CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$$

ethanol *orange* *green*

All reactants are in the breathalyzer in excess, except the ethanol, which comes from one's breath. The amount of Cr^{3+} produced is dependent on the amount of ethanol, therefore the colour change and intensity of green is dependent on the amount of ethanol provided.

How do you know the above equation is a redox equation?

$C_2H_5OH + K_2Cr_2O_7 + H_2SO_4 \Rightarrow CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$

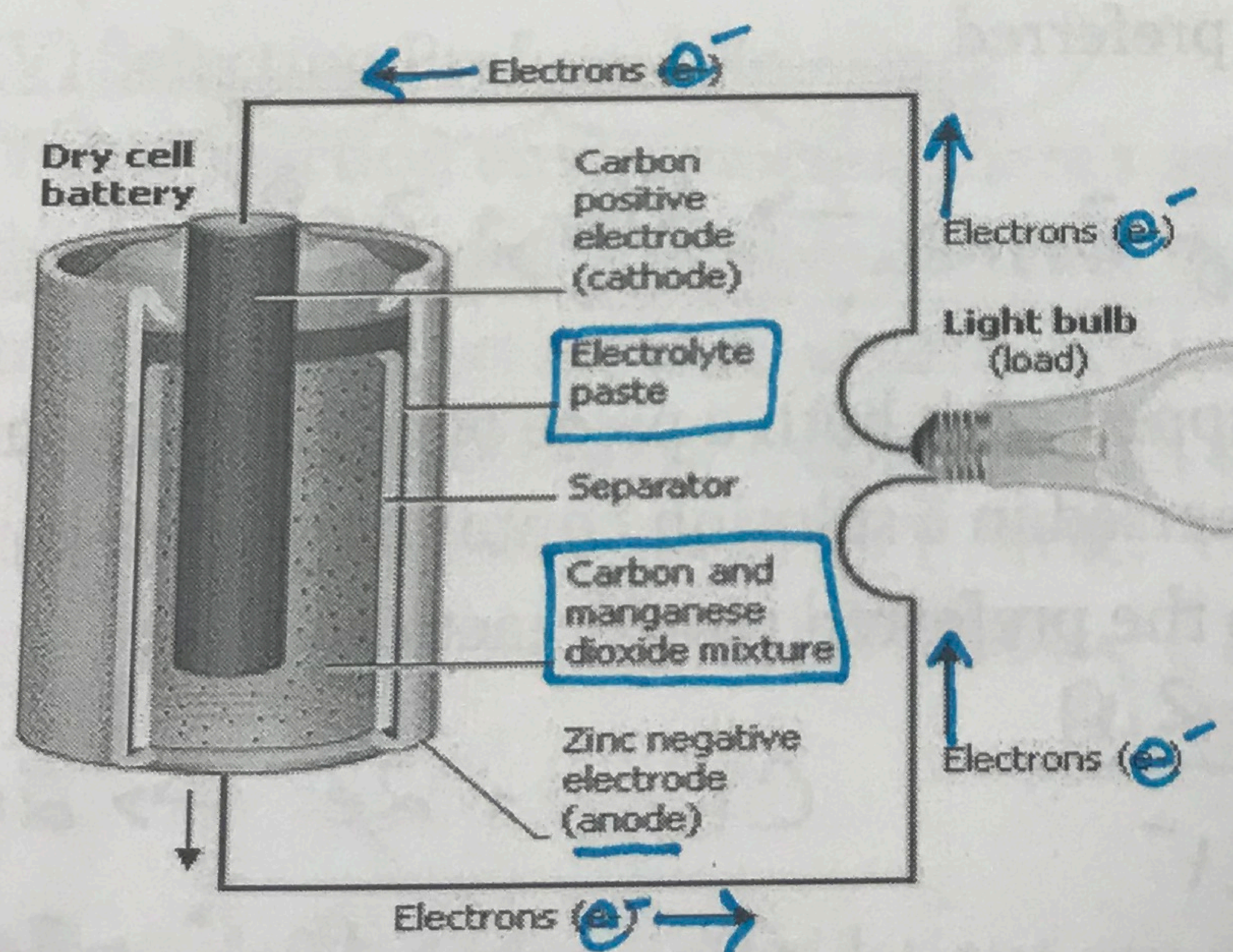
+4 +5 −2 +1 (+2) +2 +12 −14 (+6) +3 −2 −2 +1

change in oxidation numbers for C and Cr

Batteries

For any battery, electrons produced at the anode leave the battery to power a device and return via the electrical circuit to the cathode for reduction.

Zinc-Carbon Battery (aka Dry-Cell → 'Dry' because ions in a paste)



oxidation half-reaction: $Zn \Rightarrow Zn^{2+} + 2e^-$ **the Zn casing acts as the anode*

reduction half-reaction: $2MnO_2 + 2NH_4^+ + 2e^- \Rightarrow Mn_2O_3 + 2NH_3 + H_2O$

The Zn casing around the battery is the anode. A carbon rod down the centre is the inert cathode. The battery dies when all of the Zn is consumed.

Advantages: cheap materials
Disadvantages: not rechargeable, short shelf life, voltage inconsistent

KOH (basic)

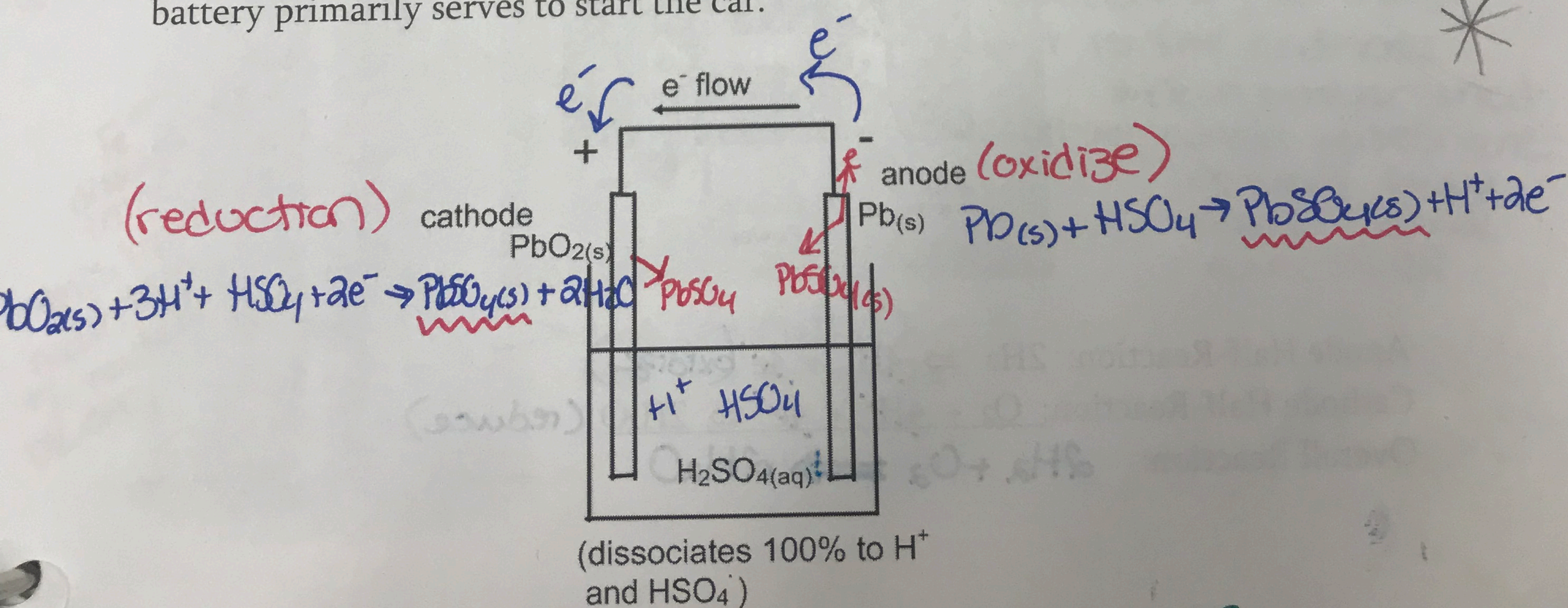
Alkaline Dry Cell (Modified Version of Zinc/Carbon Dry Cell)
Same as the zinc-carbon battery except the paste is manganese dioxide and potassium hydroxide.
The KOH electrolyte gives this battery its name (alkaline = basic). The half-reactions are the same as the zinc-carbon battery except they are under basic conditions. Alkaline batteries are the most common battery today.

Anode half-reaction: $Zn + 2OH^- \Rightarrow ZnO + H_2O + 2e^-$
Cathode Half-Reaction: $2MnO_2 + H_2O + 2e^- \Rightarrow Mn_2O_3 + 2OH^-$

Advantages: - more efficient ion transport in basic (alkaline) electrolyte
- more constant voltage than zinc/carbon battery

Disadvantages: - materials more expensive

Lead-Acid Storage Battery (Car Battery)
This is the type of battery found in automobiles. It is made up of 6 individual cells connected in series, with each cell producing 2 volts (making a 12 volt battery). Each cell consists of one Pb plate and one PbO_2 plate immersed in Sulfuric Acid. The cathode and anode do not need separate compartments since both Pb and PbO_2 are solids, thus they cannot come in direct contact. This battery primarily serves to start the car.



recharging half rxns would be the reverse of the discharging ½ rxns. (done by the alternator) req's voltage/energy

Discharging half-reactions:

Anode: $Pb_{(s)} + HSO_4^- \Rightarrow PbSO_{4(s)} + H^+ + 2e^-$

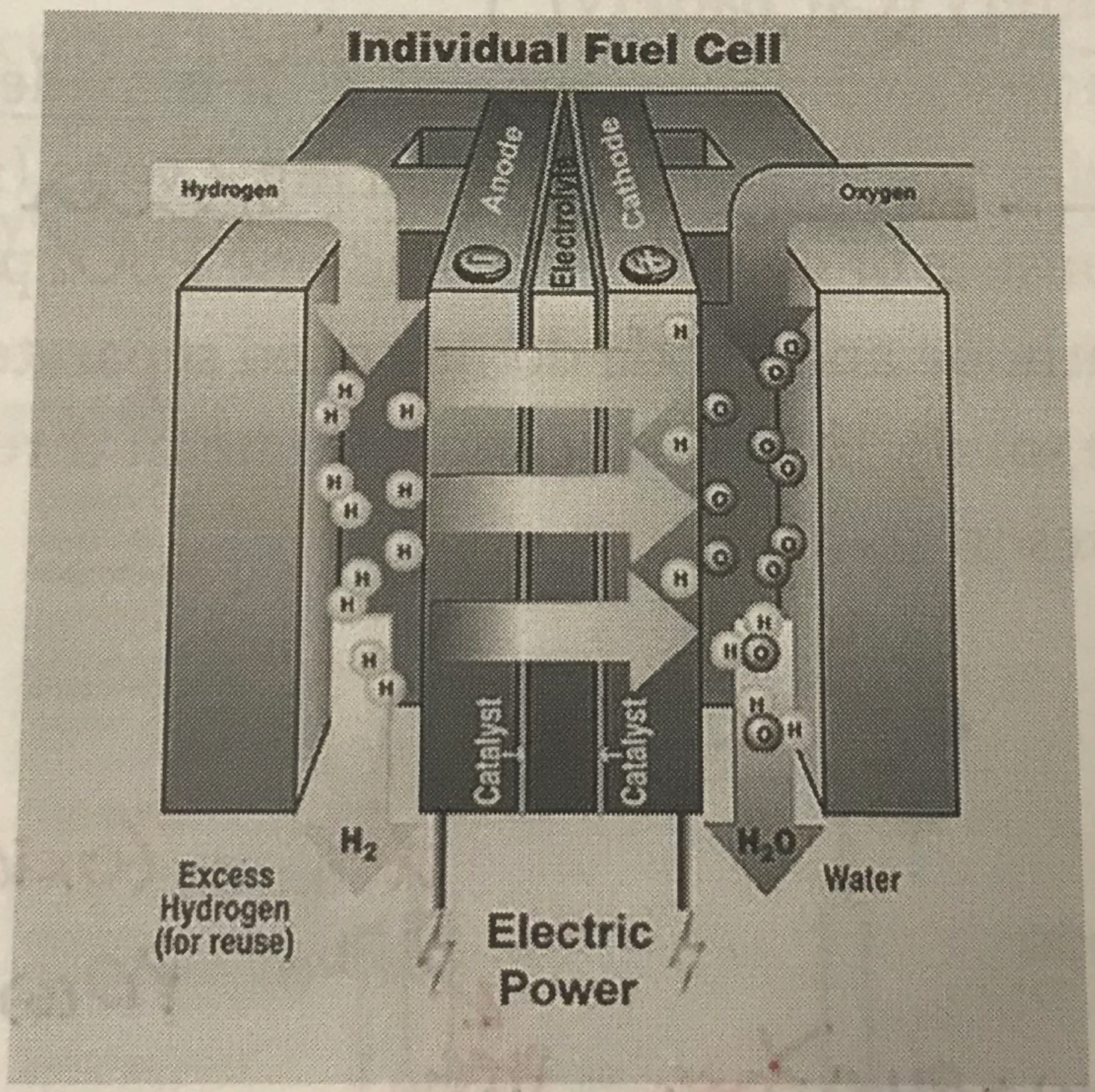
Cathode: $PbO_{2(s)} + 3H^+ + HSO_4^- + 2e^- \Rightarrow PbSO_{4(s)} + 2H_2O$

*Recharging half-reactions would be the reverse of the discharging rxns

Notice that solid $PbSO_4$ is a product of each half-reaction. It sticks to both electrodes and serves as the reactant when the battery is recharged, which is accomplished by applying a voltage to reverse the half-reactions. The *alternator* does this; it is powered by the motor to convert its mechanical energy to chemical energy, which then provides the necessary voltage to the non-spontaneous recharging reaction. Eventually, $PbSO_4$ can 'flake-off' of the electrodes, due to bumps, erratic driving etc, and fall to the bottom of the H_2SO_4 bath. This loss of reactants disallows the battery from fully recharging. Eventually, a new battery is required.

Fuel Cell

Fuel cells are different from the batteries described above because fuel cells have to be continually supplied with hydrogen and oxygen gas whereas the other batteries are self-contained units.



Anode Half-Reaction: $2H_2 \Rightarrow 4H^+ + 4e^-$ (oxidize)

Cathode Half-Reaction: $O_2 + 4H^+ + 4e^- \Rightarrow 2H_2O$ (reduce)

Overall Reaction: $2H_2 + O_2 \Rightarrow 2H_2O$