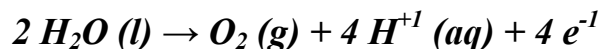


1. Water was electrolyzed, as shown in the diagram above, for 5.61 minutes using a constant current of 0.513 amper. A small amount of nonreactive electrolyte was added to the container before the electrolysis began. The temperature was 298 K and the atmospheric pressure was 1.00 atm.

(a) Write the balanced equation for the half reaction that took place at the anode.

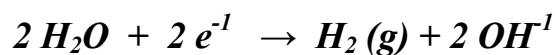
“ AN OX, RED CAT ”

(half-reaction where water is oxidized and oxygen gas is formed)



Not part of the answer (in part (a)), but occurring at the CATHODE:

(half-reaction where water is reduced and hydrogen gas is formed)



- (b) Calculate the amount of electric charge, in coulombs, that passed through the solution.

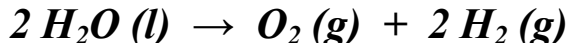
$$I = \frac{q}{t}$$

$$\frac{5.61 \text{ min}}{1 \text{ min}} \times \frac{60 \text{ sec}}{1 \text{ min}} = 336.6 \text{ sec}$$

$$q = It = (0.513 \text{ amp}) (336.6 \text{ sec}) = 173 \text{ coulombs}$$

(c) Why is the volume of O₂ (g) collected different from the volume of H₂ (g) collected, as shown in the diagram?

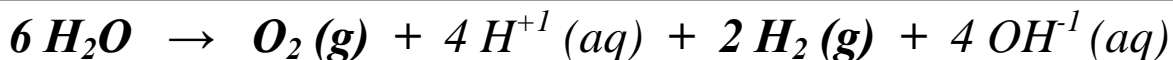
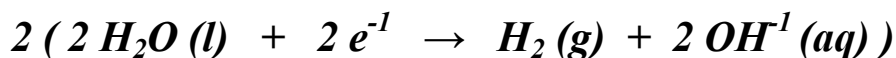
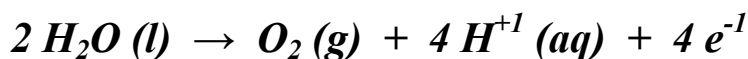
When water decomposes according to the balanced chemical equation



Twice as many moles of hydrogen are produced than moles of oxygen – recall, all gases (regardless of the mass of the molecule) behave in the same way at the same temperature and pressure.

(d) Calculate the number of moles of H₂ (g) produced during the electrolysis.

The half-reaction that takes place at the cathode is:



$$\frac{173 \text{ coulombs}}{96500 \text{ coulombs}} \times \frac{1 \text{ mol } e^{-1}}{4 \text{ mole } e^{-1}} \times \frac{2 \text{ mol H}_2}{4 \text{ mole } e^{-1}} = 8.96 \times 10^{-4} \text{ mol H}_2$$

(e) Calculate the volume, in liters, at 298 K and 1.00 atm of dry H₂ (g) produced during the electrolysis.

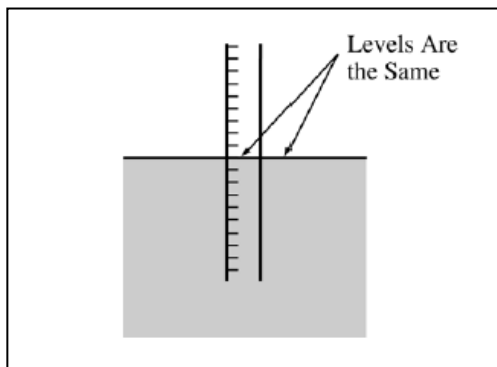
“ dry H₂ (g) “ refers to only the hydrogen gas (ignoring the water vapor usually included with the pressure of the gas BUBBLED through and COLLECTED OVER WATER)

– you TYPICALLY calculate for a dry gas – nothing new...

$$PV = nRT$$

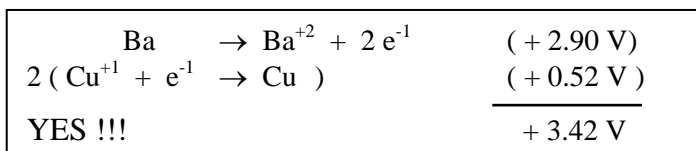
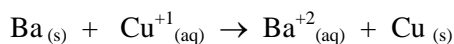
$$V = \frac{nRT}{P} = \frac{(8.96 \times 10^{-4} \text{ mol}) \left(\frac{0.0821 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (298 \text{ K})}{1 \text{ atm}} = 0.0219 \text{ L}$$

- (f) After the hydrolysis reaction was over, the vertical position of the tube containing the collected $H_2(g)$ was adjusted until the water levels inside and outside the tube were the same, as shown in the diagram below. The volume of gas in the tube was measured under these conditions of 298 K and 1.00 atm, and its volume was greater than the volume calculated in part (e). Explain.



Because the electrolysis of water occurs in water, there is some water vapor in the tube where hydrogen gas was collected. The volume calculated in part (e) was the volume of only the hydrogen gas in the tube at the given temperature and pressure. The presence of another gas (water vapor) results in greater volume at the given temperature and pressure.

2. Calculate the standard free-energy change for the reaction shown below: (Be sure to balance atoms and charges)

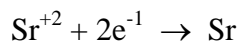


$$\Delta G = -n \mathfrak{F} E^0$$

$$\Delta G = -(2 \text{ mol } e^-) \left(\frac{96,500 \text{ J}}{V \cdot \text{mol } e^-} \right) (+ 3.42 \text{ V}) = -6.60 \times 10^5 \text{ J}$$

Negative $\Delta G = \text{spontaneous}$

3. How long will it take to plate out 3.33 kg Sr from aqueous Sr^{+2} with a current of 12.0 amps

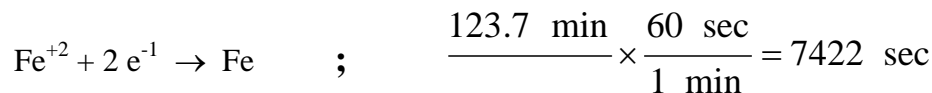


$$\frac{3.33 \text{ kg Sr}}{1.0 \text{ kg}} \times \frac{1000 \text{ g Sr}}{1 \text{ kg}} \times \frac{1 \text{ mol Sr}}{87.62 \text{ g}} \times \frac{2 \text{ mol } e^{-1}}{1 \text{ mol Sr}} \times \frac{96,500 \text{ coulombs}}{1 \text{ mol } e^{-1}} = 7.335 \times 10^7 \text{ coulombs}$$

$$I = \frac{q}{t} \quad ; \quad t = \frac{q}{I} = \frac{7.335 \times 10^7 \text{ amps} \cdot \text{sec}}{12.0 \text{ amps}} = 6.112 \times 10^5 \text{ sec}$$

$$\frac{6.112 \times 10^5 \text{ sec}}{3600 \text{ sec}} \times \frac{1 \text{ hour}}{3600 \text{ sec}} = 170 \text{ hour}$$

4. Using a current of 28.00 amps, it took 123.7 minutes to plate all of the iron from 555 mL of a solution containing Fe^{+2} .
 (a) Calculate the mass (in grams) of iron "plated out", and (b) the original concentration of Fe^{+2} in the solution?



$$q = It = (28.00 \text{ amp}) (7422 \text{ sec}) = 2.078 \times 10^5 \text{ coulombs}$$

$$\frac{2.078 \times 10^5 \text{ coulombs}}{96,500 \text{ coulombs}} \times \frac{1 \text{ mol } e^{-1}}{2 \text{ mol } e^{-1}} \times \frac{1 \text{ mol Fe}^{+2}}{2 \text{ mol } e^{-1}} = 1.08 \text{ mol}$$

$$(a) \frac{1.08 \text{ mol}}{1 \text{ mol Fe}} \times \frac{55.85 \text{ g}}{1 \text{ mol Fe}} = 60.1 \text{ g} \quad ; \quad (b) M = \frac{1.08 \text{ mol}}{0.555 \text{ L}} = 1.94 \text{ M}$$