

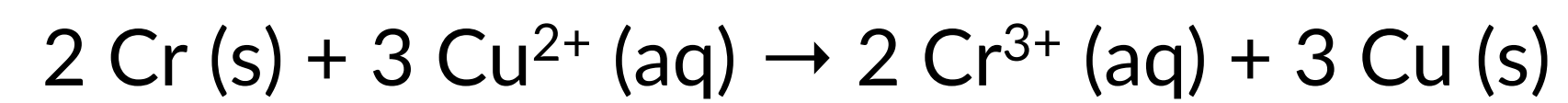


# **ELECTROCHEMISTRY**

**ELECTROCHEMICAL CELLS**

CHEMISTRY 165 // SPRING 2020

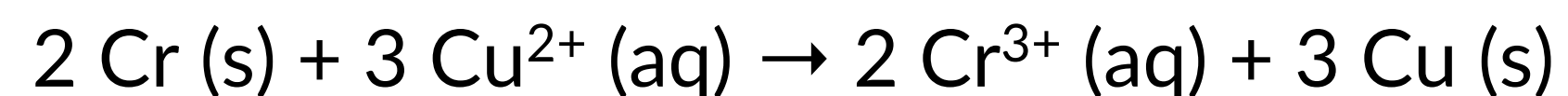
## PRACTICE PROBLEM 1



Consider a Galvanic (voltaic) cell consisting of the reaction above. Write the cell notation/diagram for this cell.

— *answer* —

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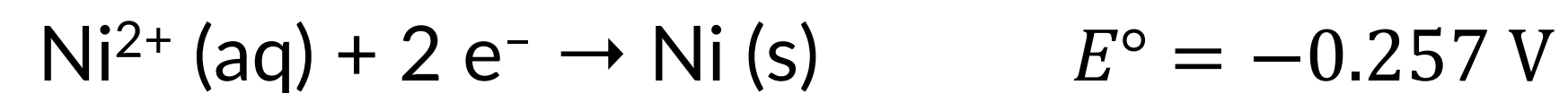
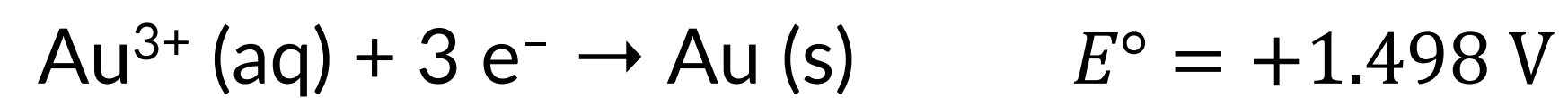
We can deconstruct the overall net ionic reaction taking place in the cell into the reduction (cathode) and oxidation (anode) reactions by considering the oxidation states of the reactants and products.



Now we can construct the cell diagram. Start with the salt bridge ( || ) which separates the anode and cathode. The anode is on the left and cathode is on the right in the diagram. To the left/right of the salt bridge, we write the aqueous solutions (Cr<sup>3+</sup> and Cu<sup>2+</sup>). Finally, we write the electrodes (Cr and Cu) on the far ends of the diagram, which is separated from the aqueous solutions with a phase-separating line ( | ).



## PRACTICE PROBLEM 2

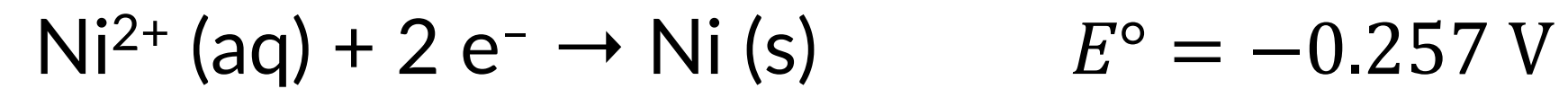
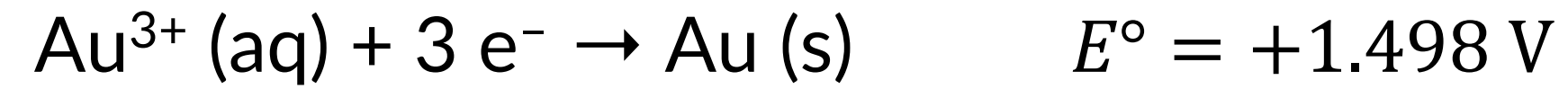


(a) What is the standard cell potential for a Galvanic (voltaic) cell consisting of the two half-reactions above?

(b) Write the cell diagram.

— *answer* —

## PRACTICE PROBLEM 2



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— answer —

(a) Recognize that a Galvanic (voltaic) cell is one that is **spontaneous**, so  $\Delta G_{\text{cell}}^{\circ} < 0$  and  $E_{\text{cell}}^{\circ} > 0$  because  $\Delta G_{\text{cell}}^{\circ} = -nFE_{\text{cell}}^{\circ}$ .

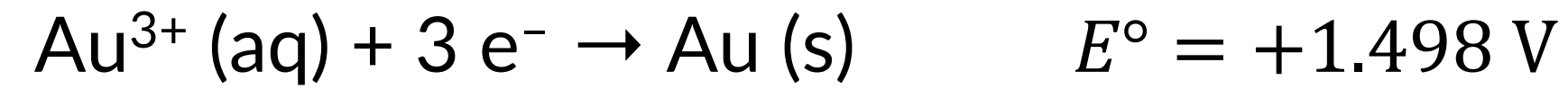
Generally, the reaction with the more positive value of  $E^{\circ}$  will be the reduction reaction (cathode)—remember that the  $E$  value is related to the spontaneity ( $\Delta G$ ) of the reaction, so we want to keep the more spontaneous reduction reaction as the cathode.

This means that the other half-reaction is the oxidation reaction (anode).

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = +1.498 \text{ V} - (-0.257 \text{ V}) = +1.755 \text{ V}$$



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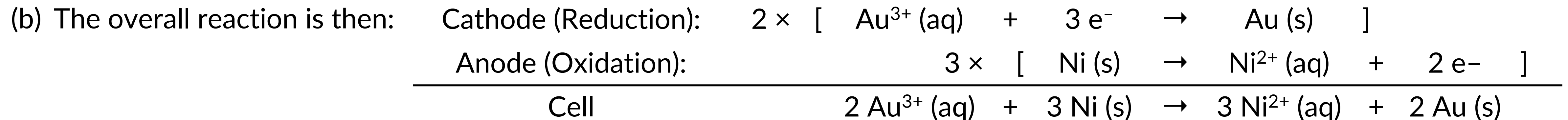
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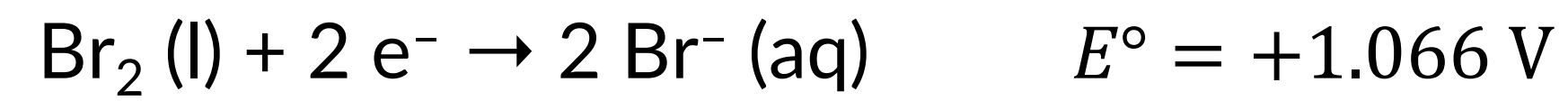
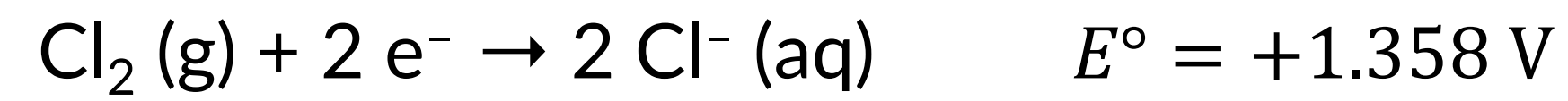
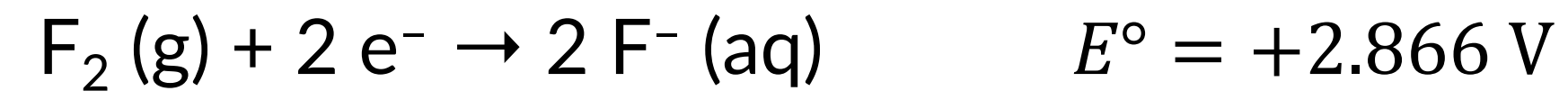


And the cell diagram is:  $\text{Ni} (\text{s}) \mid \text{Ni}^{2+} (1.0 \text{ M}) \parallel \text{Au}^{3+} (1.0 \text{ M}) \mid \text{Au} (\text{s})$

Recall that standard conditions for electrochemistry are:  $T = 298.15 \text{ K}$ ;  $P = 1.0 \text{ atm}$ ;  $[ ] = 1.0 \text{ M}$

## PRACTICE PROBLEM 3

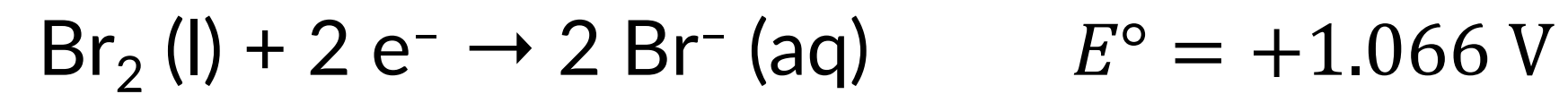
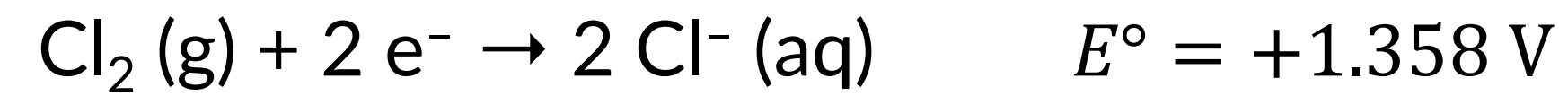
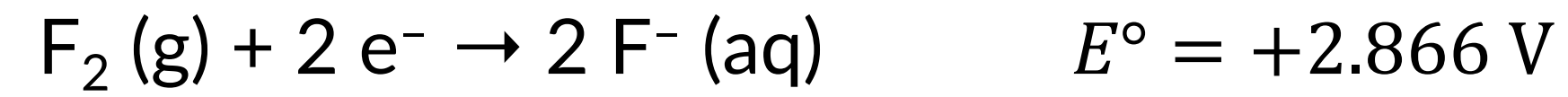
Based on the following standard reduction potentials, which of the following is the strongest oxidizing agent?



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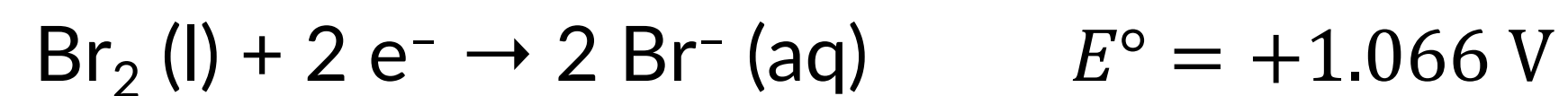
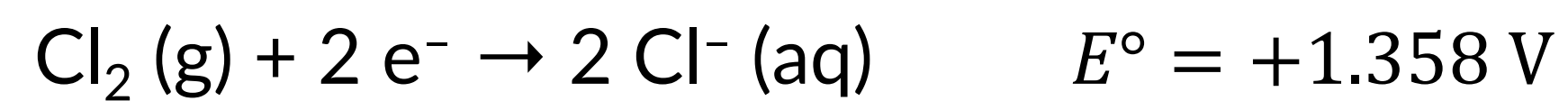
An oxidizing agent is a species that is reduced.

Recall that the larger the  $E^\circ$  value, the more spontaneous/favorable the reduction reaction.

Therefore,  $\text{F}_2 (\text{g})$  is the strongest oxidizing agent.



## PRACTICE PROBLEM 4

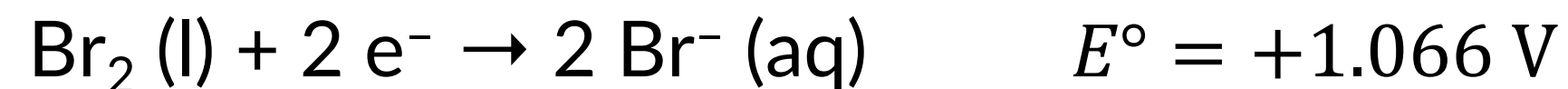
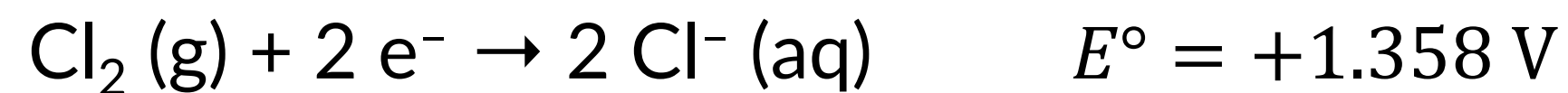


(a) Write the cell diagram for a Galvanic (voltaic) cell constructed at standard conditions from the two reactions above.

(b) Calculate the standard Gibbs free energy change of the cell ( $\Delta G_{\text{cell}}^\circ$ ).

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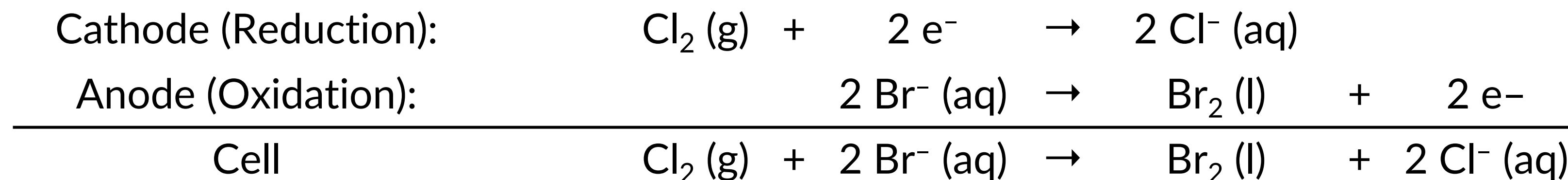


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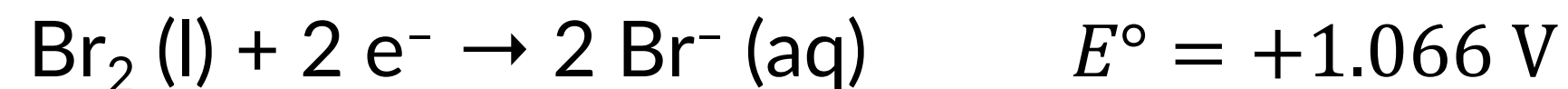
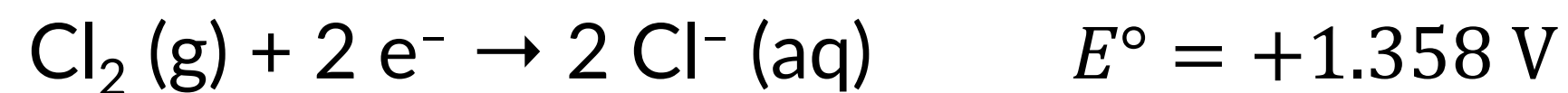


Therefore, the cell diagram is: **Pt (s) | Br<sub>2</sub> (l) | Br<sup>-</sup> (1.0 M) || Cl<sup>-</sup> (1.0 M) | Cl<sub>2</sub> (1.0 atm) | Pt (s)**

Recall that standard conditions for electrochemistry are:  $T = 298.15 \text{ K}$ ;  $P = 1.0 \text{ atm}$ ;  $[ ] = 1.0 \text{ M}$

In this case, we need to use an inert solid electrode (Pt) because Br<sub>2</sub> and Cl<sub>2</sub> are not solid. In addition, we need to use an addition phase-separating line ( | ) to separate the aqueous phases (Br<sup>-</sup> and Cl<sup>-</sup>) from the others (Br<sub>2</sub> and Cl<sub>2</sub>).

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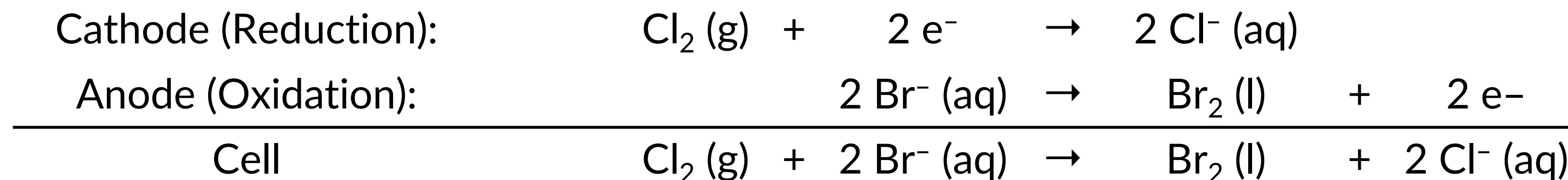


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(b) The standard cell potential is  $E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ = +1.358 \text{ V} - 1.066 \text{ V} = +0.292 \text{ V}$ .

The standard Gibbs free energy change can now be calculated using  $1 \text{ V} = 1 \text{ J/C}$ :

$$\Delta G_{\text{cell}}^\circ = -nFE_{\text{cell}}^\circ = -(2 \text{ mol } \text{e}^-) \left( 96500 \frac{\text{C}}{\text{mol } \text{e}^-} \right) \left( 0.292 \frac{\text{J}}{\text{C}} \right) = -56400 \text{ J} = -56.4 \text{ kJ}$$