Shown below are entries from the standard potentials table in the back of your book:

\[
2 \text{H}^+ (aq) + 2 e^- \rightarrow \text{H}_2 (g) \quad 0.000 \text{ v}
\]

\[
\text{Cu}^{2+} (aq) + 2 e^- \rightarrow \text{Cu} (s) \quad 0.337
\]

(3) 1. Write the half-reactions and a balanced equation for the reaction of H\(^+\) ions and copper metal.

\[
2\text{H}^+(aq) + 2e^- \rightarrow \text{H}_2(g) \quad 0.000 \text{ V}
\]

\[
\text{Cu} (s) \rightarrow \text{Cu}^{2+}(aq) + 2e^- \quad -0.337
\]

\[
2\text{H}^+(aq) + \text{Cu}(s) + 2e^- \rightarrow \text{Cu}^{2+}(aq) + \text{H}_2(g) + 2e^-
\]

(2) 2. What is oxidized? \(\text{Cu}\)  
What is reduced? \(\text{H}^+\)  
What is the oxidizing agent? \(\text{H}^+\)  
What is the reducing agent? \(\text{Cu}\)

(1) 3. What is the cell potential for this reaction under standard electrochemical conditions?

\[
\varepsilon^\circ_{\text{cell}} = -0.337 \text{ V}
\]

(2) 4. Is this reaction spontaneous under standard electrochemical conditions? Why or why not?

No. A negative value for \(\varepsilon\) indicates the reaction is not spontaneous (or spontaneous in the reverse direction).

(2) 5. How would increasing the concentration of H\(^+\) ions effect the potential for the reaction? Explain.

\(\text{H}^+\) is a reactant, therefore it would make \(\varepsilon\) more positive (or less negative) since \(Q = \frac{[\text{Cu}^{2+}][\text{H}_2]}{[\text{H}^+]^2}\)