·reduction potential: tendency for a chemical species to be reduced)
·the reference value is the reduction of hydrogen, assigned a value of 0 V
·the more positive the reduction potential, the greater the tendency of reduction
standard cell potentials : $E°_{cell} = E°_{red}(\text{cathode}) - E°_{red}(\text{anode})$
·multiplying half-reactions by integers to balance the equation does NOT change the E° (energy per electron) of material)
·reduction potentials are intensive: do not depend on amount)

standard Reduction Potential Example: $2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$ (1M, 1M)
·standard hydrogen electrode: reduction + 0V by convention
what is the reduction potential for this reaction, given info shown?
·$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$, where volt-meter reading = +0.34
·copper is cathode w/ hydrogen, and is more readily reduces,
→ means that cell potential is positive

Question 2: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$, spontaneous rxn, positive cell potential
- Hydrogen reduced more readily, so cell potential = negative
- Reduction potential of Zn^{2+}, $E°_{red} = 0.76\,V$
- Reduction potential of Cu^{2+}, $E°_{red} = +0.34\,V$
- For this voltaic cell: $E°_{cell} = +0.34\,V - (-0.76\,V) = 1.1\,V$

convention: anode on the left

Question 3: Cu shown as the anode
what does volt meter read? spontaneous rxn - positive #
$E°_{red}(Cu) = 0.34\,V$
$E°_{red}(Ag) = 0.80\,V$
$E°_{cell} = E°_{red}(\text{cathode}) - E°_{red}(\text{anode})$
$= 0.80\,V - 0.34\,V = \boxed{0.46\,V}$

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Top of table: strong oxidizing agents
Bottom : strong reducing agents
·spontaneous redox reactions produce a positive $E°_{cell}$ + negative $\Delta G°$
$E°_{cell} = E°_{red}(\text{cathode}) - E°_{red}(\text{anode})$
· Relationship between cell potential and ΔG : $\Delta G° = -nFE°_{cell}$
· F : Faraday constant