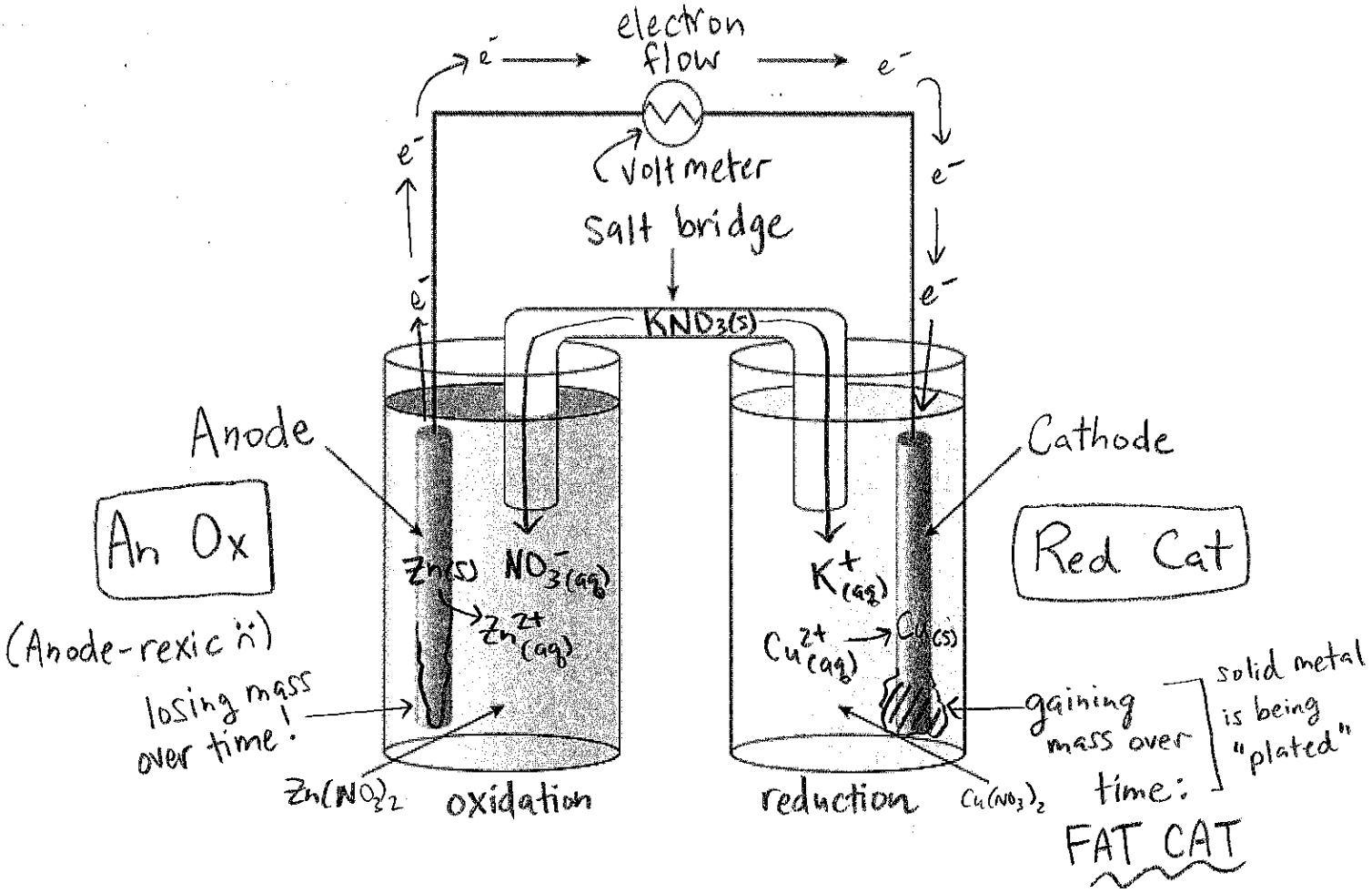
Galvanic/ ~~Voltaic~~ 26 Cells

- Redox reactions involve the transfer of electrons from one substance to another, and thus have the potential to generate an electric current (i.e. flow of electrons).
- To use that current, we need to Separate the place where oxidation is occurring from the place where reduction is occurring.
    - Current is the number of electrons that flow through the system per second.
    - Current is measured in amperes, or Amp (A) = 1 coulomb of charge per second.
- This known as a **voltaic** (or **galvanic**) cell: the most common form of which is a battery!
- Galvanic (voltaic) cells are always thermodynamically favorable (spontaneous) and thus have a + E°_{cell}.



Parts of the Galvanic Cell

1. anode (–): the electrode where **oxidation** occurs (loses mass as reaction progresses)
2. Cathode (+): the electrode where **reduction** occurs (gains mass as reaction progresses, metal 'plated')
3. Salt bridge (or disk): bridge between cells whose purpose is to provide ions to balance the charge and complete the circuit.
    - **An**ions (–) flow to the anode; **cat**ions (+) flow to the Cathode
    - If the salt bridge is removed, current will slow and then Stop (V = 0) as charge builds up in the half-cells.
4. Voltmeter: measures the cell potential (emf or E°) in volts
    - Over time, voltage in the galvanic cell will decrease as [reactants] ↓ and [products] ↑

Important Notes about Galvanic Cells NOT salt bridge! ~~25~~ 27 F.A.T. cathode

1. **Electron Flow**: ALWAYS through the wire from anode to cathode (alpha order)
2. **Ion-Ion or Ion-Gas Redox**: a voltaic cell can be constructed where the underlying redox reaction involves a gas or the conversion from one ion to another (unlikely a traditional voltaic cell in which reduction is from ion to solid and oxidation is from solid to ion).
   - ❖ Requires an inert electrode: this doesn't take part in the redox reaction but provides a surface on which the electrons can transfer; commonly made of platinum (expensive) or graphite (cheap)

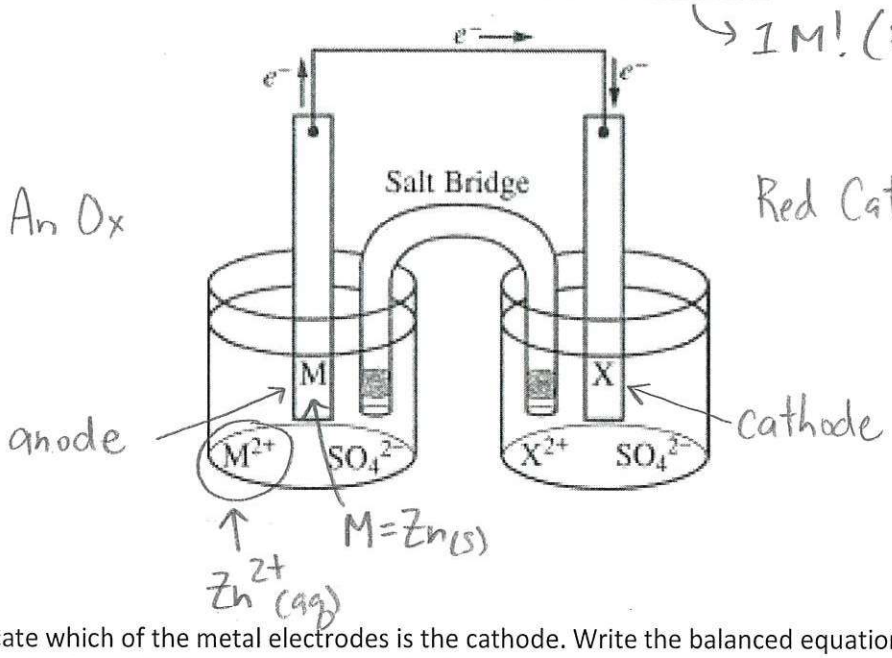
Let's Practice!

1. The diagram below shows the experimental setup for a typical electrochemical cell that contains two standard half-cells. The cell operates according to the reaction represented by the following equation. (2002 #7, form B)

$$\underset{\text{ox}}{Zn(s)} + \underset{\text{red}}{Ni^{2+}(aq)} \rightarrow Zn^{2+}(aq) + Ni(s)$$

   a. Identify M and M^{2+} in the diagram and specify the initial concentration for M^{2+} in solution.

→ 1 M! (standard cell)



   b. Indicate which of the metal electrodes is the cathode. Write the balanced equation for the reaction that occurs in the half-cell containing the cathode.

Cathode = metal X = $Ni_{(s)}$

$$Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)}$$

   c. Describe what would happen to the cell voltage if the salt bridge was removed. Explain.

Cell voltage would drop to zero over time if the salt bridge was removed, b/c there would be no ions to balance out the positive charge build-up in the anode and negative charge build-up in the cathode.