The standard dry cell (battery) has the following shorthand notation:

\[ \text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{MnO}_2(\text{s}), \text{Mn}_2\text{O}_3(\text{s}) \mid \text{graphite(s)} \]

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Which component of the battery is the anode? **Zn(s) electrode**

Give the balanced half equation that takes place at the anode.

**Oxidation at the anode:** \( \text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2e^- \)

Which component of the battery is the cathode? **Graphite electrode**

Give the balanced half equation that takes place at the cathode.

**Reduction at the cathode:** \( 2\text{MnO}_2(\text{s}) + 2\text{H}^+(\text{aq}) + 2e^- \rightarrow \text{Mn}_2\text{O}_3 + \text{H}_2\text{O}(\text{l}) \)
The bacterium *Azotobacter chroococcum*, growing aerobically in a medium free of nitrogen containing compounds, obtains all of its nitrogen by the "fixation" of atmospheric N\(_2\). The solubility of N\(_2\) in water is governed by the following equilibrium:

\[
N_2(aq) \rightleftharpoons N_2(g) \quad K = 1.6 \times 10^3 \text{ atm L mol}^{-1}
\]

What is the concentration of dissolved N\(_2\) available to the bacterium at 1.0 atm and 30 °C? (Air is 78% N\(_2\).)

The equilibrium constant is given by

\[K = \frac{[N_2(g)]}{[N_2(aq)]}\]

where \(N_2(g)\) is the partial pressure of N\(_2\) (g).

The atmospheric pressure is the sum of the partial pressures of the constituent gases in the air. As N\(_2\) represents 78% of the air and the total pressure is 1.0 atm, the partial pressure of N\(_2\) is given by

\[p_{N_2}(g) = 0.78 \times 1.0 \text{ atm} = 0.78 \text{ atm}\]

Hence, the concentration of dissolved N\(_2\) is:

\[\frac{[N_2(aq)]}{K} = \frac{0.78 \text{ atm}}{1.6 \times 10^3 \text{ atm L mol}^{-1}} = 4.9 \times 10^{-4} \text{ M}\]

Answer: 4.9 \times 10^{-4} \text{ M}

A culture of these bacteria (1.0 L) grows to a density of 0.84 mg dry weight per mL of culture and has a nitrogen content of 7.0% of the dry weight. What volume of air at 1.0 atm and 30 °C would supply this nitrogen requirement?

As the density of the culture is 0.84 mg mL\(^{-1}\), the mass of 1.0 L (1000 mL) is 1000 mL \times (0.86 \times 10^{-3} \text{ g mL}^{-1}) = 0.86 g. The nitrogen content is 7.0% so the mass of nitrogen in the culture is 0.07 \times 0.86 g = 0.060 g.

As nitrogen has a atomic mass of 14.01 g mol\(^{-1}\), this mass corresponds to:

\[
\text{moles of nitrogen atoms} = \frac{\text{mass of nitrogen}}{\text{atomic mass}} = \frac{0.060 \text{ g}}{14.01 \text{ g mol}^{-1}} = 0.0043 \text{ mol}
\]

Nitrogen is present in the air as N\(_2\) so the number of moles of N\(_2\) required is \(\frac{1}{2} \times 0.0043 \text{ mol} = 0.0021 \text{ mol}\).

The partial pressure due to N\(_2\) is 0.78 atm if the atmospheric pressure is 1.0 atm. Using \(PV = nRT\), the volume of air containing 0.0021 mol of N\(_2\) (g) is:

\[
V = \frac{nRT}{P} = \frac{(0.021 \text{ mol}) \times (0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}) \times ((30+273) \text{ K})}{0.78 \text{ atm}}
\]

\[= 0.068 \text{ L} = 68 \text{ mL}\]

Answer: 0.068 L or 68 mL
The mechanism of copper toxicity to aquatic organisms is unknown. Most theories attribute the toxicity to the Cu\(^{2+}\) species because Cu\(^+\) is unstable in aqueous solution. Given the half-reactions and half-cell potentials on the data page, show that it is electrochemically favourable for Cu\(^+\)(aq) to react with itself to form Cu\(^{2+}\)(aq) and Cu(s).

The reaction of interest is:

\[ 2\text{Cu}^+(aq) \rightarrow \text{Cu(s)} + \text{Cu}^{2+}(aq) \]

It can be considered as a reduction of 2Cu\(^+\)(aq) to make 2Cu(s), followed by the oxidation of one of the Cu(s) to make Cu\(^{2+}\)(aq).

The half cell reactions and potentials are:

- \[ \text{Cu}^+(aq) + e^- \rightarrow \text{Cu(s)} \quad E^\circ = +0.53 \text{ V} \]
- \[ \text{Cu(s)} \rightarrow \text{Cu}^{2+}(aq) + 2e^- \quad E^\circ = -0.34 \text{ V} \text{ (reversed as oxidation required)} \]

Hence the reaction:

\[ 2\text{Cu}^+(aq) \rightarrow \text{Cu(s)} + \text{Cu}^{2+}(aq) \]

has a cell potential \( E^\circ = ((+0.53) + (-0.34)) \text{ V} = +0.19 \text{ V} \)

As \( E^\circ > 0 \), the reaction is spontaneous

The Co\(^{3+}\) ion is unstable in aqueous solution, but for a different reason to Cu\(^+\) above. Using the table of reduction potentials on the data page, propose the reason why this might be so.

The Co\(^{3+}\) / Co\(^{2+}\) half cell is:

\[ \text{Co}^{3+}(aq) + e^- \rightarrow \text{Co}^{2+}(aq) \quad E^\circ = +1.82 \text{ V} \]

This is sufficiently positive to be able to oxidise water:

\[ 2\text{H}_2\text{O}(l) \rightarrow \text{O}_2(g) + 4\text{H}^+(aq) + 4e^- \quad E^\circ = -1.23 \text{ V} \]

The overall reaction is:

\[ 4\text{Co}^{3+} + 2\text{H}_2\text{O}(l) \rightarrow 4\text{Co}^{2+}(aq) + \text{O}_2(g) + 4\text{H}^+(aq) \]

The electrode potential is \( E^\circ = ((+1.82) + (-1.23)) \text{ V} = +0.59 \text{ V} \)

As \( E^\circ > 0 \), the reaction is spontaneous