**Rules for Understanding Redox Reactions**

1. By convention, redox half-reactions are written as reductions:

$$\text{oxidant} + e^- \rightarrow \text{reductant}$$

eg. $O_2 + 4e^- (+ 4\ H^+) \rightarrow 2H_2O$ $E_0 = +0.82$ V

2. The more positive the $E_0$ value, the higher the affinity of the oxidant for $e^-$s; ie the better an oxidant it is. The more negative the $E_0$ value, the better reductant it is – it gives up $e^-$s more easily.

3. Under standard conditions, an oxidant (eg. $O_2$, $E_0 = +0.82$ V) will oxidize a reductant (eg. NADH, $E_0 = -0.32$ V) if the reductant exhibits a more negative $E_0$ value. Conversely, a reductant will reduce any oxidant whose $E_0$ value is more positive. Thus $O_2$ will oxidize NADH (NADH will reduce $O_2$) under standard conditions.

4. The sign is reversed for half-reactions written in the opposite direction.

5. A net positive $E_0$ value for any conjugate redox pair (eg. $O_2$/NADH) means the forward reaction is favorable under standard conditions. A net negative $E_0$ value means the reverse reaction is favorable. (opposite of $\Delta G^0$)

Eg. $\frac{1}{2}\ O_2 + NADH + H^+ \rightarrow H_2O + NAD^+$ $E^o = +1.14$ V

Thus, oxidation of NADH by oxygen is highly favorable.