

Oxidation Numbers and Redox Reactions

Answer key with worked solutions

A. Understanding

1. Definitions.

Oxidation is the loss of electrons; the oxidation number **increases**. **Reduction** is the gain of electrons; the oxidation number **decreases**.

B. Oxidation numbers

2. Oxidation numbers.

(a) -3 (b) +5 (c) +6 (d) +6 (e) +3 (f) -1 (peroxide) (g) +4 (h) +5. For example, (h): $x + 3(-2) = -1$, so $x = +5$.

C. Agents and half-equations

3. Chlorine and bromide.

Br⁻ is oxidized ($-1 \rightarrow 0$) and **Cl₂ is reduced** ($0 \rightarrow -1$). **Oxidizing agent: Cl₂; reducing agent: Br⁻.**

4. Half-equations.

Oxidation: $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$. Reduction: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$.

5. Redox or not?

(a) **Not redox**: no oxidation number changes. (b) **Redox** (disproportionation): O goes from -1 to -2 in H₂O and to 0 in O₂. (c) **Not redox**: a precipitation in which all oxidation numbers stay the same.

D. Balancing

6. Copper and silver.

Cu loses 2e^- ; each Ag^+ gains 1e^- , so 2Ag^+ are needed: $\text{Cu} + 2\text{Ag}^+ \longrightarrow \text{Cu}^{2+} + 2\text{Ag}$.
Charge: $+2 = +2$. ✓

7. Dichromate and iron(II).

Reduction: $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \longrightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$

Oxidation ($\times 6$): $6\text{Fe}^{2+} \longrightarrow 6\text{Fe}^{3+} + 6\text{e}^-$

Overall: $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{Fe}^{2+} \longrightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + 6\text{Fe}^{3+}$. Charge: $-2 + 14 + 12 = +24$ on the left; $+6 + 18 = +24$ on the right. ✓

8. Copper and nitrate.

Oxidation ($\times 3$): $3\text{Cu} \longrightarrow 3\text{Cu}^{2+} + 6\text{e}^-$

Reduction ($\times 2$): $2\text{NO}_3^- + 8\text{H}^+ + 6\text{e}^- \longrightarrow 2\text{NO} + 4\text{H}_2\text{O}$

Overall: $3\text{Cu} + 2\text{NO}_3^- + 8\text{H}^+ \longrightarrow 3\text{Cu}^{2+} + 2\text{NO} + 4\text{H}_2\text{O}$. Charge: $-2 + 8 = +6 = +6$. ✓

E. Challenge

9. Permanganate.

(a) Reduction ($\times 2$): $2\text{MnO}_4^- + 16\text{H}^+ + 10\text{e}^- \longrightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O}$; oxidation ($\times 5$):
 $5\text{C}_2\text{O}_4^{2-} \longrightarrow 10\text{CO}_2 + 10\text{e}^-$

Overall: $2\text{MnO}_4^- + 5\text{C}_2\text{O}_4^{2-} + 16\text{H}^+ \longrightarrow 2\text{Mn}^{2+} + 10\text{CO}_2 + 8\text{H}_2\text{O}$. Charge: $-2 - 10 + 16 = +4 = +4$. ✓

(b) $n(\text{MnO}_4^-) = 0.0200\text{ mol/L} \times 0.02000\text{ L} = 4.00 \times 10^{-4}\text{ mol}$

$n(\text{Fe}^{2+}) = 4.00 \times 10^{-4}\text{ mol MnO}_4^- \times \frac{5\text{ mol Fe}^{2+}}{1\text{ mol MnO}_4^-} = 2.00 \times 10^{-3}\text{ mol}$

$[\text{Fe}^{2+}] = \frac{2.00 \times 10^{-3}\text{ mol}}{0.02500\text{ L}} = 0.0800\text{ M}$