

Oxidation-Reduction (Redox) Reactions

- ☞ A reaction in which one species transfers electrons to another is an *oxidation-reduction reaction*, also called a *redox reaction*.

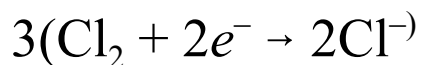


- ✓ **Oxidation** is the *loss* of electrons by a substance.
- ✓ **Reduction** is the *gain* of electrons by a substance.



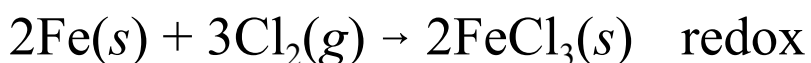
Fe⁰ "pushes" e's ☞

oxidation



Cl₂ "pulls" e's ☞

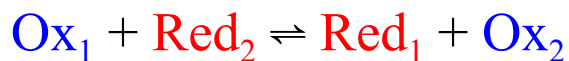
reduction



- ☞ ***There is never an oxidation without a reduction, and vice versa!***

Oxidizing Agents and Reducing Agents

- ✓ An *oxidizing agent* (or *oxidant*) is a substance that causes another substance to be oxidized and is itself reduced.
- ✓ A *reducing agent* (or *reductant*) is a substance that causes another substance to be reduced and is itself oxidized.
- ☞ In these terms, all redox reactions take on the general form

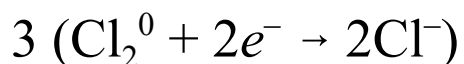


Oxidation Numbers and Redox

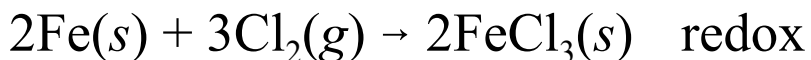
- ▲ When a species is **oxidized**, one of its atoms goes to a higher (more positive or less negative) oxidation number.
- ▼ When a species is **reduced**, one of its atoms goes to a lower (less positive or more negative) oxidation number.



Fe oxidation number
increases \Rightarrow **oxidation**



Cl oxidation number
decreases \Rightarrow **reduction**



Balancing Redox Equations by the Ion-Electron Method

- 1. Separate the skeletal equation into two half reactions.** Each half reaction refers to the conversion of a species in either its oxidized or reduced form into a related species in either its reduced or oxidized form. One half reaction will be a **reduction** and the other will be an **oxidation**.
- 2. Balance each half reaction separately.** Balance atoms on each side of a half reaction by inspection, using H_2O , H^+ (if in acid), or OH^- (if in base) to make the balance in hydrogen and/or oxygen, if needed. **Do not add any other new species (e.g., O_2 , H_2) unless already a part of the skeletal half reaction.**
- 3. Balance the net charge across each half reaction by adding electrons to the side with the more positive net ionic charge.** If by this process electrons are added on the left side of a half reaction, the half reaction is a **reduction**. If electrons are added to the right side, the half reaction is an **oxidation**. **(If you add electrons to the same side in both half reactions, something is wrong!)**
- 4. Multiply both half-reactions by appropriate factors (usually whole numbers), so that the number of electrons is the same in both half reactions and will cancel when the two are added together.**
- 5. Add the two multiplied half reactions together to obtain the overall redox equation.**
- 6. Check the balance.** **No electrons should appear in the overall redox equation.** Not only should there be a balance in atoms across the equation, but also the net charge on both sides of the equation should be equal.

Work-Around Technique for Difficult Basic Cases

- ☹ Balancing H and O in basic redox reactions sometimes can be difficult, because both OH^- and H_2O contain both elements.
- ☺ A trick to balance troublesome basic cases:
 - ✓ Balance any troublesome half-reaction or the entire redox reaction as if it were in acid first.
 - ✓ Then add equal numbers of OH^- to both sides of the acid-balanced equation to "neutralize" any H^+ to become H_2O ; i.e., $\text{H}^+ + \text{OH}^- = \text{H}_2\text{O}$.

Example:

