

## Oxidation-Reduction (Redox) Reactions

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

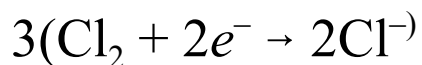


- ✓ **Oxidation** is the *loss* of electrons by a substance.
- ✓ **Reduction** is the *gain* of electrons by a substance.



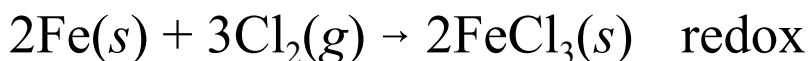
Fe<sup>0</sup> "pushes" e's ☞

**oxidation**



Cl<sub>2</sub> "pulls" e's ☞

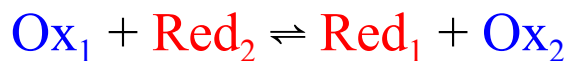
**reduction**



- ☞ ***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

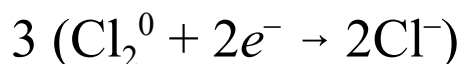


## Oxidation Numbers and Redox

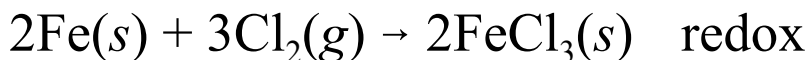
- ▲ When a species is **oxidized**, one of its atoms goes to a higher (more positive or less negative) oxidation number.
- ▼ When a species is **reduced**, one of its atoms goes to a lower (less positive or more negative) oxidation number.



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  $\text{H}_2\text{O}$ ,  $\text{H}^+$  (if in acid), or  $\text{OH}^-$  (if in base) to make the balance in hydrogen and/or oxygen, if needed. **Do not add any other new species (e.g.,  $\text{O}_2$ ,  $\text{H}_2$ ) 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  $\text{OH}^-$  and  $\text{H}_2\text{O}$  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  $\text{OH}^-$  to both sides of the acid-balanced equation to "neutralize" any  $\text{H}^+$  to become  $\text{H}_2\text{O}$ ; i.e.,  $\text{H}^+ + \text{OH}^- = \text{H}_2\text{O}$ .

Example:

