# Questions Answered by Oxidation Numbers

| | |
|---|---|
| Is the reaction redox? | If any atoms change their oxidation number, yes. |
| What's oxidized? | The element that increases its oxidation number |
| What's reduced? | The element that decreases its oxidation number |
| What's the reducing agent? | The substance with the element oxidized |
| What's the oxidizing agent? | The substance with the element reduced |

# Steps for Determination of Oxidation Numbers

- **Step 1:** Assign oxidation numbers to as many atoms as you can using the guidelines described on the next slide.
- **Step 2:** To determine oxidation numbers for atoms not described on the next slide, use the following guideline.
  - The sum of the oxidation numbers for each atom in the formula is equal to the overall charge on the formula. (This includes uncharged formulas where the sum of the oxidation numbers is zero.)

# Oxidation Numbers for Common Elements

| | Oxidation number | Examples | Exceptions |
|---|---|---|---|
| **Pure element** | 0 | The oxidation number for each atom in Zn, $H_2$, and $S_8$ is zero. | none |
| **Monatomic ions** | charge on ion | Cd in $CdCl_2$ is +2.<br>Cl in $CdCl_2$ is –1. | none |
| **Fluorine in the combined form** | –1 | F in $AlF_3$ is –1.<br>F in $CF_4$ is –1. | none |
| **Oxygen in the combined form** | –2 | O in ZnO is –2.<br>O in $H_2O$ is –2. | O is –1 in peroxides, such as $H_2O_2$ |
| **Hydrogen in the combined form** | +1 | H in $H_2O$ is +1.<br>H in $HSO_4^-$ is +1. | H is –1 when combined with a metal. |

# Examples

- Determine the oxidation number for each of the atoms or ions in the following formulas.
- **$N_2$**
  0 (Uncharged element)
- **$N^{3-}$**
  –3 (monatomic ion)
- **$H_2S$**
  +1 for H (covalently bonded H)
  –2 for S (solve $\boldsymbol{0 = 2(1) + x}$ for x)
- **LiH**
  +1 for Li, –1 for H (monatomic ions)

# Examples

- Determine the oxidation number for each of the atoms or ions in the following formulas.
- **$K_2HPO_4$**

  +1 for K (monatomic ion)

  +1 for H (covalently bonded in polyatomic ion)

  –2 for O (combined oxygen, not a peroxide)

  +5 for P (solve the following equation for x)

  $$0 = 2(1) + 1 + x + 4(-2)$$
  $$0 = 3 + x - 8$$
  $$0 = x - 5$$
  $$+5 = x$$

# Examples

- Determine the oxidation number for each of the atoms or ions in the following formulas.
- **$HCr_2O_7^-$**

+1 for H (covalently bonded in a polyatomic ion)

–2 for O (combined oxygen, not a peroxide)

+6 for Cr (solve the following equation for x)

$$-1 = 1 + 2x + 7(-2)$$
$$-1 = 1 + 2x - 14$$
$$-1 = 2x - 13$$
$$12 = 2x$$
$$6 = x$$

# Examples

- Determine the oxidation number for each of the atoms or ions in the following formulas.
- **$NH_4^+$**

  +1 for H (covalently bonded in polyatomic ion)

  –3 for N (solve the following equation for x)

  $+1 = x + 4(+1)$

  $+1 = x + 4$

  $-3 = x$

# Examples

- Determine the oxidation number for each of the atoms or ions in the following formulas.
- **$H_2PO_2^-$**

  +1 for H (covalently bonded in polyatomic ion)

  –2 for O (combined oxygen)

  +1 for P (solve the following equation for x)

  $$-1 = 2(1) + x + 2(-2)$$
  $$-1 = 2 + x - 4$$
  $$-1 = x - 2$$
  $$+1 = x$$

# Examples

- Determine the oxidation number for each of the atoms or ions in the following formulas.
- **$Cu_2SO_4$**

  +1 for Cu (monatomic ion)

  –2 for O (combined oxygen, not a peroxide)

  +6 for S (solve the following equation for x)

  $$0 = 2(1) + x + 4(-2)$$
  $$0 = 2 + x - 8$$
  $$0 = x - 6$$
  $$+6 = x$$

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| What's the oxidizing agent? | The substance with the element reduced |

# Examples

- In one part of the steel manufacturing process, carbon is combined with iron to form pig iron. Pig iron is easier to work with than pure iron because it has a lower melting point (about 1130 °C, compared to 1539 °C for pure iron) and is more pliable. The following equations describe its formation. Determine the oxidation number for each atom in the formulas. Decide whether each reaction is a redox reaction, and if it is, identify what is oxidized, what is reduced, what the oxidizing agent is, and what the reducing agent is.

$$2C(s) + O_2(g) \rightarrow 2CO(g)$$

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$$

$$2CO(g) \rightarrow C(\text{in iron}) + CO_2(g)$$

# Exercise

$$\overset{0}{2C(s)} + \overset{0}{O_2(g)} \rightarrow 2\overset{+2\,-2}{CO}(g)$$

- Oxidation numbers

  0 for C and O in $O_2$ (uncharged elements)

  –2 for O in CO (combined oxygen, not a peroxide)

  +2 for C in CO (solve 0 = x + (-2) for x)

- C goes from 0 to +2 so it is oxidized.
- O goes from 0 to -2 so it is reduced
- O in $O_2$ is reduced so $O_2$ is the oxidizing agent.
- C in C(s) is oxidized so C is the reducing agent.

# Exercise

$$\overset{+3\ -2}{Fe_2O_3(s)} + \overset{+2\ -2}{CO(g)} \rightarrow \overset{0}{2Fe(l)} + \overset{+4\ -2}{3CO_2(g)}$$

- Oxidation numbers
  - –2 for O in all formulas with O (combined oxygen, not a peroxide)
  - +3 for Fe in $Fe_2O_3$ (solve **$0 = 2x + 3(-2)$** for x)
  - +2 for C in CO (solve **$0 = x + (-2)$** for x)
  - 0 for Fe in Fe(l) (uncharged element)
  - +4 for C in $CO_2$ (solve **$0 = x + 2(-2)$** for x)
- Fe goes from +3 to 0, so it is reduced.
- C goes from +2 to +4, so it is oxidized
- Fe in $Fe_2O_3$ is reduced, so $Fe_2O_3$ is the oxidizing agent.
- C in CO is oxidized, so CO is the reducing agent.

# Exercise

$$\overset{+2\ -2}{2CO(g)} \rightarrow \overset{0}{C(\text{in iron})} + \overset{+4\ -2}{CO_2(g)}$$

- Oxidation numbers

  –2 for O in all formulas with O (combined oxygen, not a peroxide)

  +2 for C in CO (solve **0 = x + (-2)** for x)

  0 for C (uncharged element)

  +4 for C in $CO_2$ (solve **0 = x + 2(-2)** for x)

- C in CO goes from +2 to 0 in C, so it is reduced.
- C in CO goes from +2 to +4 in $CO_2$, so it is oxidized.
- C in CO is both oxidized and reduced, so CO is both the oxidizing agent and the reducing agent. (This is called disproportionation.)