

Name:

Date:

Period:

Seat #:

Show all work

Assigning oxidation numbers:

Determine the oxidation number of the underlined element.

- | | | | | | | | |
|-----|---------------------------|-----|---------------------------|-----|--|-----|---|
| 1. | <u>Be</u> Cl ₂ | 2. | N <u>O</u> | 3. | Na ₂ <u>S</u> O ₃ | 4. | H ₂ <u>O</u> ₂ |
| 5. | Ag <u>Br</u> | 6. | <u>Au</u> Cl ₃ | 7. | H <u>N</u> O ₃ | 8. | H ₂ <u>Sn</u> O ₃ |
| 9. | <u>S</u> O ₃ | 10. | <u>U</u> F ₆ | 11. | Ba <u>Cr</u> O ₄ | 12. | Ca <u>Se</u> O ₄ |
| 13. | H <u>I</u> | 14. | H ₂ <u>Se</u> | 15. | K ₂ <u>Pt</u> Cl ₆ | 16. | <u>Ni</u> SO ₄ |
| 17. | <u>N</u> H ₃ | 18. | H <u>Cl</u> O | 19. | <u>N</u> H ₄ Cl | 20. | (NH ₄) ₂ <u>Te</u> |

- | | | | | | | | |
|-----|-------|-----|-------|-----|-------|-----|-------|
| 1. | _____ | 2. | _____ | 3. | _____ | 4. | _____ |
| 5. | _____ | 6. | _____ | 7. | _____ | 8. | _____ |
| 9. | _____ | 10. | _____ | 11. | _____ | 12. | _____ |
| 13. | _____ | 14. | _____ | 15. | _____ | 16. | _____ |
| 17. | _____ | 18. | _____ | 19. | _____ | 20. | _____ |

The Half-Reaction Method:

- Write the equation as two half-reactions. Include the particles (atoms, ions, molecules) that are involved in change of oxidation state.
- Balance each half-reaction with respect to atoms and charges; first atoms other than H and O, then O with H₂O and H with H⁺, and ionic charges with electrons (e⁻).
- Equalize the number of electrons lost in the oxidation half-reaction with the number of electrons gained in the reduction half-reaction.
- Add the two half-reactions to form a balanced net ionic equation.

[a]	
$\text{HCl} + \text{K}_2\text{Cr}_2\text{O}_7 \rightarrow \text{KCl} + \text{CrCl}_3 + \text{Cl}_2$	
14, 1, 2, 2, 3, 7H ₂ O	
Reduction half-reaction	Oxidation half-reaction
Overall:	

[b]	
$\text{FeCl}_2 + \text{KMnO}_4 + \text{HCl} \rightarrow \text{FeCl}_3 + \text{KCl} + \text{MnCl}_2 + \text{H}_2\text{O}$	
5, 1, 8, 5, 1, 1, 4H ₂ O	
Reduction half-reaction	Oxidation half-reaction
Overall:	

[c]	
$\text{CuS} + \text{NO}_3^- \rightarrow \text{Cu}^{2+} + \text{S} + \text{NO}$	
$3, 2, 8\text{H}^+, 3, 3, 2, 4\text{H}_2\text{O}$	
Reduction half-reaction	Oxidation half-reaction
Overall:	

Balance the following redox reactions in acidic solutions:

[d] $\text{HNO}_3 + \text{S} \rightarrow \text{NO}_2 + \text{H}_2\text{SO}_4 + \text{H}_2\text{O}$
[e] $\text{KMnO}_4 + \text{HCl} + \text{H}_2\text{S} \rightarrow \text{KCl} + \text{MnCl}_2 + \text{S}$
[f] $\text{FeCl}_3 + \text{H}_2\text{S} \rightarrow \text{FeCl}_2 + \text{HCl} + \text{S}$
[g] $\text{Cu} + \text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{NO}_2$