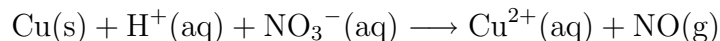


Balancing Redox Equations

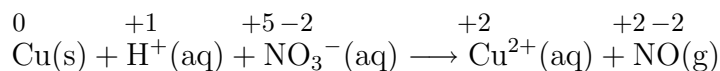
Strategy for redox equations in acidic solutions

Problem: When dilute nitric acid is poured on a piece of copper metal, copper(II) ions and the gas nitric oxide, NO, are formed. Write the balanced equation for the reaction.

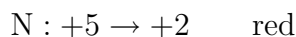
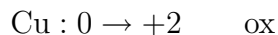
Step 0: Write the “skeleton equation”



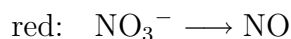
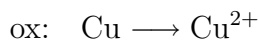
and determine the oxidation state (the oxidation state is *per atom*!)



So we have

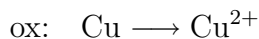


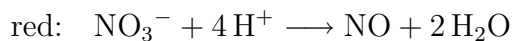
Step 1: Write the “skeleton” half-reactions



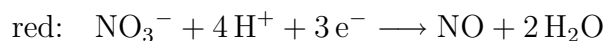
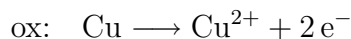
Step 2: Balance each half-reaction “atomically”

- all atoms other than H and O (You can use *any* of the species that appear in the skeleton equation—Step 0—for this purpose.)
- balance O atoms by adding H₂O
- balance H atoms by adding H⁺



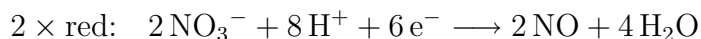


Step 3: *Balance the electric charges by adding electrons*

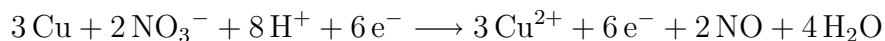


The electrons have to appear on the *right* hand side for the *oxidation* half-reaction, and on the *left* hand side for the *reduction* half-reaction.

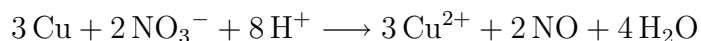
Step 4: *Prepare the two half-equations for summation by making the number of electrons the same in both, i.e, find the least common multiple*



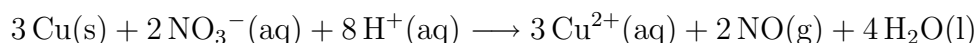
Step 5: *Combine the two half-reactions*



Step 6: *Simplify*



Step 7: *Indicate the state of each species*



This is the fully balanced net ionic equation.

Check by writing a table:

species	left	right
Cu	3	3
N	2	2
O	6	6
H	8	8
charge	+6	+6

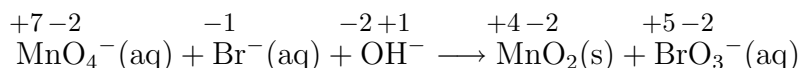
*Strategy for redox equations in **basic** solutions*

In basic solutions, no H^+ are available to balance H!

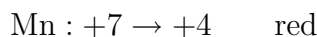
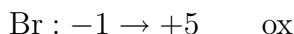
Strategy: Pretend the solution is acidic and carry out a “neutralization reaction” at the end.

Problem: The reaction of permanganate ions with bromide ions in basic solution yields solid manganese(IV) oxide and bromate ions. Write the fully balanced net equation.

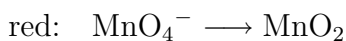
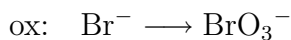
Step 0: Write the “skeleton equation” and determine the oxidation state (the oxidation state is *per* atom!)



So we have

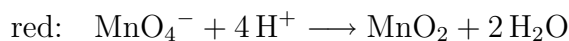
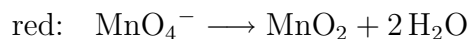
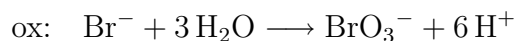
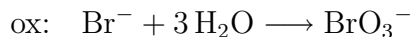


Step 1: Write the “skeleton” half-reactions

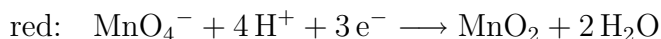
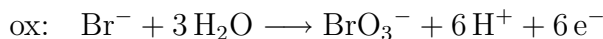


Step 2: Balance each half-reaction “atomically”

- all atoms other than H and O (You can use *any* of the species that appear in the skeleton equation—Step 0—for this purpose.)
- balance O atoms by adding H_2O
- balance H atoms by adding H^+

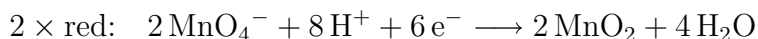
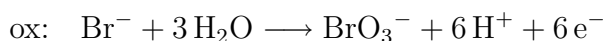


Step 3: *Balance the electric charges by adding electrons*

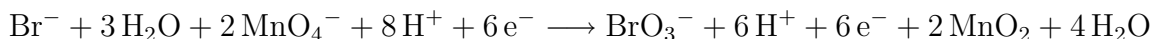


The electrons have to appear on the *right* hand side for the *oxidation* half-reaction, and on the *left* hand side for the *reduction* half-reaction.

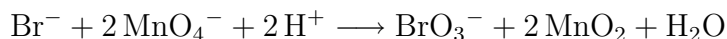
Step 4: *Prepare the two half-equations for summation by making the number of electrons the same in both, i.e, find the least common multiple*



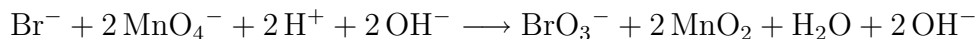
Step 5: *Combine the two half-reactions*



Step 6: *Simplify*



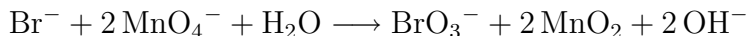
Step 6a: *Change to **basic** solution by adding as many OH^- to **both** sides as there are H^+*



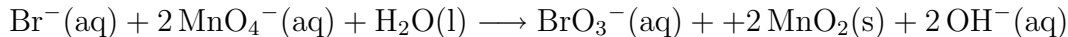
“neutralization”: Combine the H^+ and the OH^- to form H_2O



simplify



Step 7: *Indicate the state of each species*



This is the fully balanced net ionic equation.

Check by writing a table:

species	left	right
Br	1	1
Mn	2	2
O	9	9
H	2	2
charge	-3	-3