

BIG PICTURE!!!

How does Electricity Flow?

charged particles (electrons in a metal
OR
ions in solution)
move through a substance

→ Movable Charges from an ELECTRICAL SOURCE (BATTERY)

Electrolytic Cell

↳ electrical outlet

↳ charges (e^-) flow when
you give it a power source

→ Movable Charges from a CHEMICAL SOURCE (SOLUTION)

Voltaic Cell

↳ Ions in Sol'n

↳ charges (ions) flow b/c sol'n
contains the ions

* runs on its own until ions run out!

Redox Reaction Reminders:

Complete the following questions on your own with no notes or help from friends.
After each part, check your answers. If you score lower than an 80% (if you miss more than 1 question) on any part, see Ms. Monaghan for a mini-lesson to rehab that topic.

Part 1-Assigning Oxidation Numbers

- What is the formula of titanium (II) oxide?
☒ a. TiO b. Ti₂O c. TiO₂ d. Ti₂O₃
- What is the oxidation state of chlorine in NaCl?
☒ a. -1 b. +1 c. +3 d. +5
- What is the oxidation state of nitrogen in NaNO₂?
a. +1 ☒ b. +3 c. +2 d. +4
- What is the oxidation number of chromium in the chromate ion, CrO₄²⁻?
☒ a. +6 b. +3 c. +2 d. +8
- Given the following reaction: $\text{Zn(s)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{ZnSO}_4\text{(aq)} + \text{Cu(s)}$ The oxidation number of Zn changes from:
☒ a. 0 to +2 b. +2 to 0 c. 0 to -2 d. -2 to 0

Part 2-Identifying Redox Reactions

- All chemical reactions have a conservation of
a. mass, only c. mass and charge, only
b. charge and energy, only ☒ d. mass, charge, and energy
- Which equation shows conservation of both mass and charge?
☒ a. $\text{Cl}_2 + \text{Br}^- \rightarrow \text{Cl}^- + \text{Br}_2$ c. $\text{Cu} + 2 \text{Ag}^+ \rightarrow \text{Cu}^{2+} + \text{Ag}$
☒ b. $\text{Zn} + \text{Cr}^{3+} \rightarrow \text{Zn}^{2+} + \text{Cr}$ ☒ d. $\text{Ni} + \text{Pb}^{2+} \rightarrow \text{Ni}^{2+} + \text{Pb}$
- Given the balanced ionic equation: $2\text{Al(s)} + 3\text{Cu}^{2+}\text{(aq)} \rightarrow 2\text{Al}^{3+}\text{(aq)} + 3\text{Cu(s)}$
Compared to the total charge of the reactants, the total charge of the products is
a. Less b. Greater ☒ c. The same
- During which process does an atom gain one or more electrons?
a. Transmutation ☒ c. Reduction
b. Oxidation d. Neutralization
- Which reaction is an example of an oxidation reduction reaction?
a. $\text{AgNO}_3 + \text{KI} \rightarrow \text{AgI} + \text{KNO}_3$
☒ b. $\text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu(NO}_3)_2 + 2 \text{Ag}$
c. $2 \text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + 2 \text{H}_2\text{O}$
d. $\text{Ba(OH)}_2 + 2 \text{HCl} \rightarrow \text{BaCl}_2 + 2 \text{H}_2\text{O}$

DR can NEVER be Redox!

Part 3- Identifying Species Oxidized and Reduced

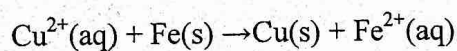
- When a neutral atom undergoes oxidation, the atom's oxidation state
 - decreases as it gains electrons
 - increases as it gains electrons
 - decreases as it loses electrons
 - ☒ increases as it loses electrons
- Which change in oxidation number indicates oxidation?
 - ☒ -1 to +2
 - 1 to -2
 - +2 to -3
 - +3 to +2
- Which changes occur when Pt^{2+} is reduced?
 - The Pt^{2+} gains electrons and its oxidation number increases.
 - ☒ The Pt^{2+} gains electrons and its oxidation number decreases.
 - The Pt^{2+} loses electrons and its oxidation number increases.
 - The Pt^{2+} loses electrons and its oxidation number decreases.
- Given the equation: $\text{C(s)} + \text{H}_2\text{O(g)} \rightarrow \text{CO(g)} + \text{H}_2\text{(g)}$ Which species undergoes reduction?
 - C(s)
 - C^{2+}
 - ☒ H^+
 - $\text{H}_2\text{(g)}$
- Given the equation: $3\text{Au(s)} + 2\text{Fe}^{+3}\text{(aq)} \rightarrow 3\text{Au}^{+2}\text{(aq)} + 2\text{Fe(s)}$
 - Gold is reduced as it loses electrons
 - ☒ Iron is reduced as it gains electrons
 - Gold is oxidized as it loses electrons
 - Iron is Oxidized as it gains electrons

Part 4- Putting it All Together

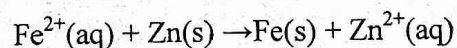
Base your answers to questions 1 through 4 on the information below.

In a laboratory investigation, a student constructs a voltaic cell with iron and copper electrodes. Another student constructs a voltaic cell with zinc and iron electrodes. Testing the cells during operation enables the students to write the balanced ionic equations below.

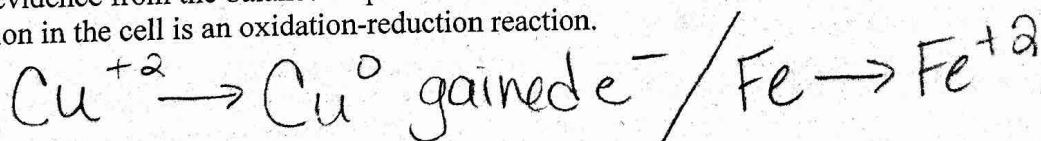
Cell with iron and copper electrodes:



Cell with zinc and iron electrodes:



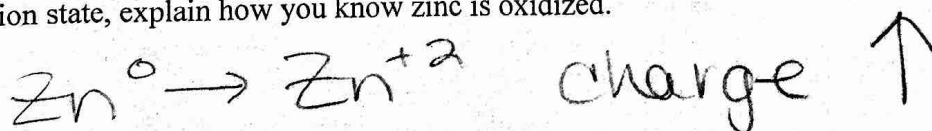
- State evidence from the balanced equation for the cell with iron and copper electrodes that indicates the reaction in the cell is an oxidation-reduction reaction.



- Identify the particles transferred between Fe^{2+} and Zn during the reaction in the cell with zinc and iron electrodes.

electrons

- In terms of oxidation state, explain how you know zinc is oxidized.



- In terms of electrons lost or gained, explain how you know zinc is oxidized.

electrons lost.

NOTES

Redox Reminders

Redox = Oxidation-Reduction

↳ Both have to happen!

LEO

→ oxidation

→ lose e^-

→ charge \uparrow (less neg)

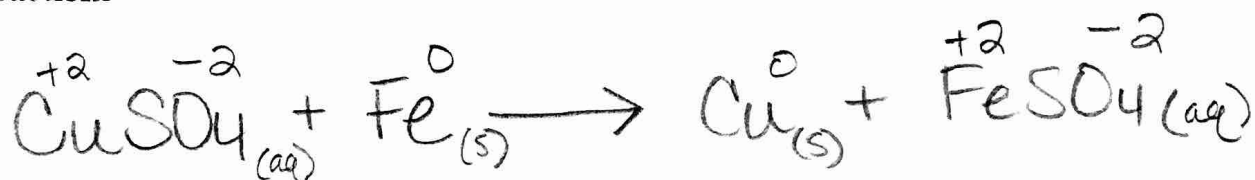
GER

→ reduction

→ gain e^-

→ charge \downarrow (more neg)

Half-Reactions



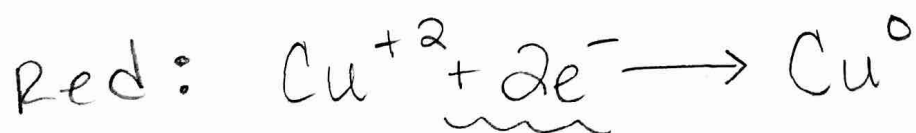
① Assign Oxidation #'s (charges)

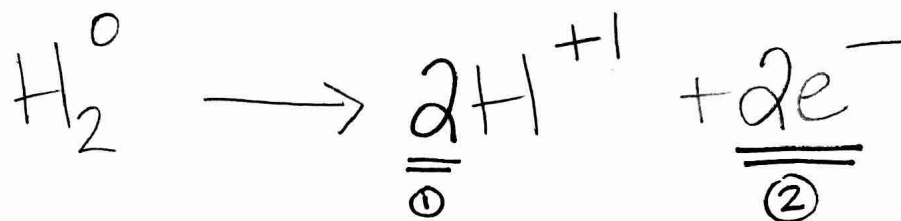
② ID who was reduced/oxidized

Red: Cu^{+2}

Ox: Fe^0

③ Write $1/2$ Rxns → show the electrons
→ make charges balance

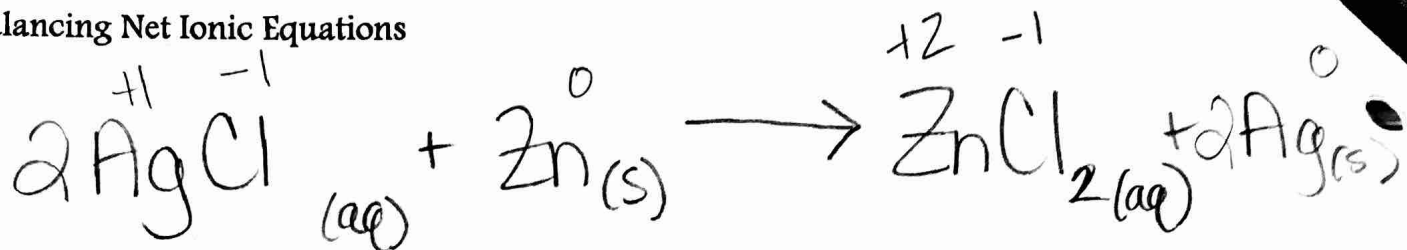




- ① Balance for mass \rightarrow coefficients
- ② Balance for charge \rightarrow add e^-

NOTES

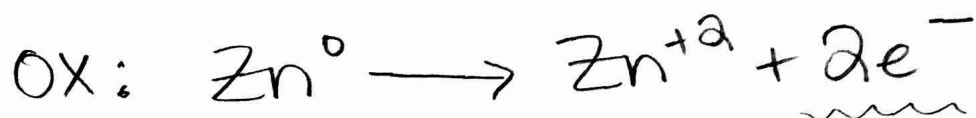
Balancing Net Ionic Equations



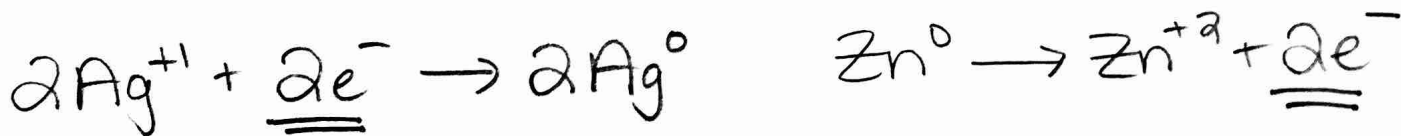
① Assign oxidation #'s

② Redox? who ox/red?

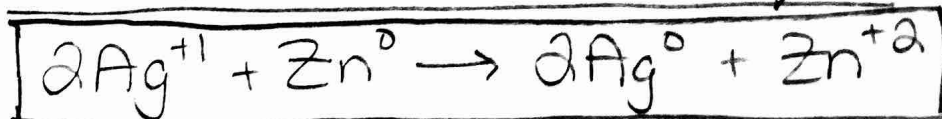
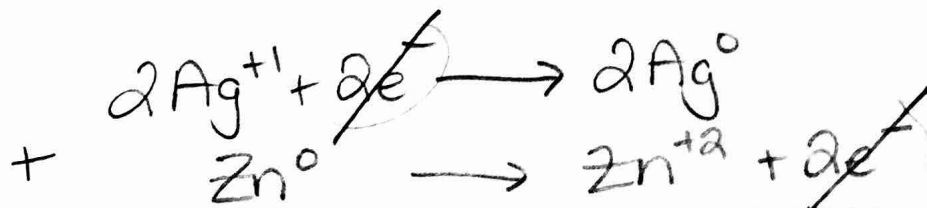
③ Write $\frac{1}{2}$ rxns



④ multiply $\frac{1}{2}$ rxns so $\# e^- \text{ lost} = \# e^- \text{ gained}$



⑤ add $\frac{1}{2}$ rxns
to cancel out
 e^- and end
w/ net ionic
EQ



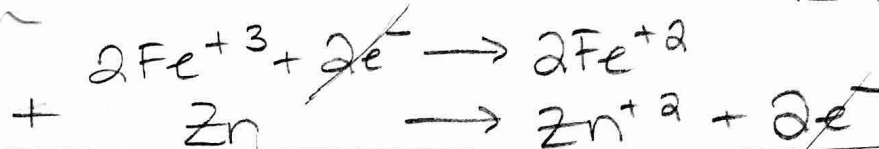
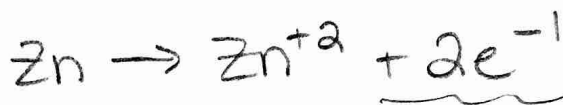
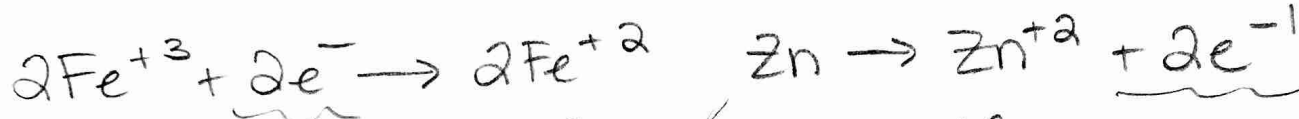
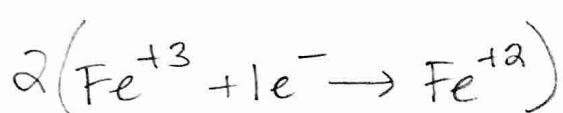
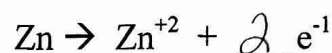
net
ionic
equation
14

Part 1:

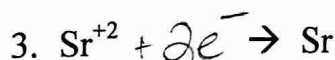
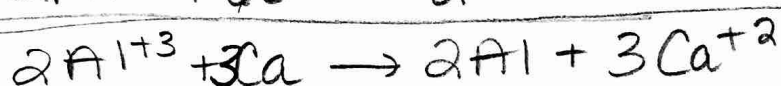
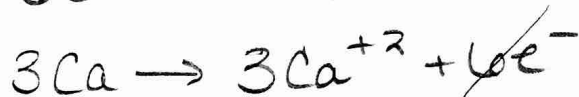
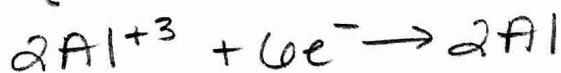
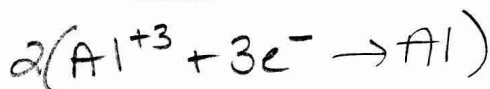
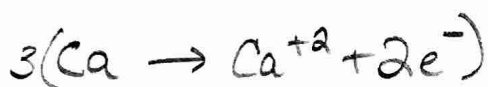
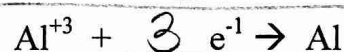
- ☐ Balance the half reactions in each pairing using the correct number of electrons.
- ☐ Identify which is the oxidation and which is the reduction half reaction.
- ☐ Combine the half reactions in order to produce the **balanced** net ionic equation.
- ☐ Explain or show how charge is conserved



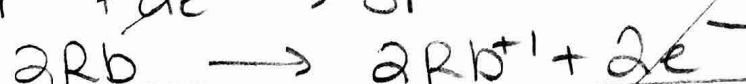
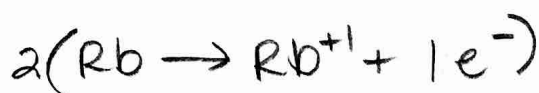
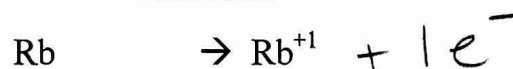
and



and

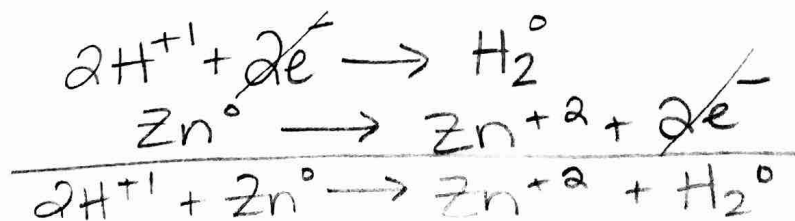
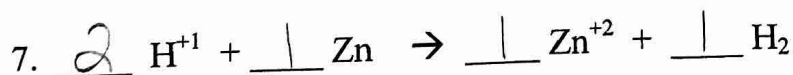
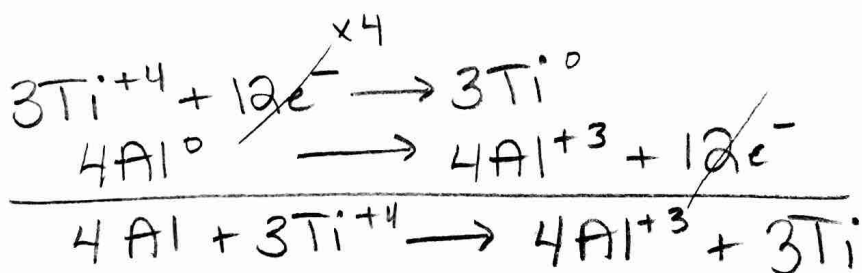
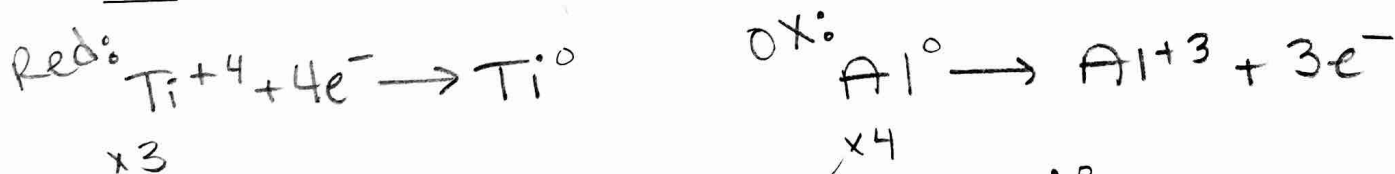
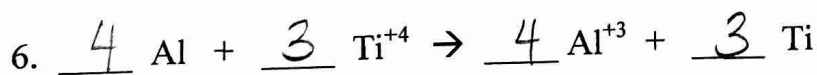
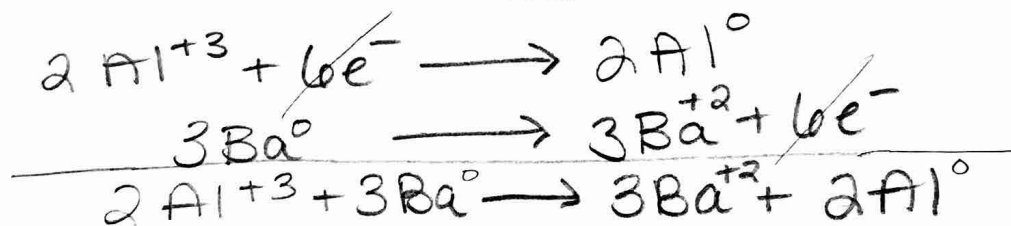
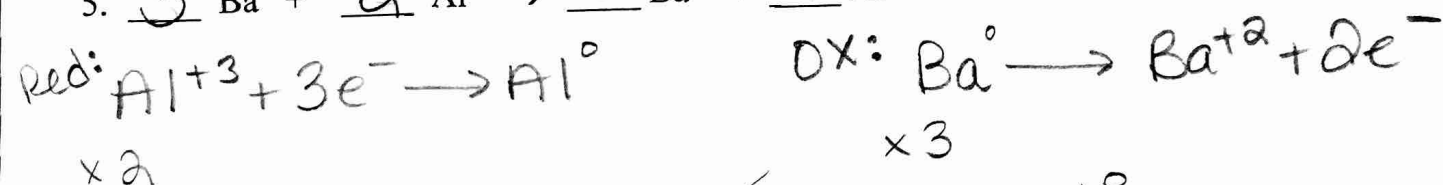
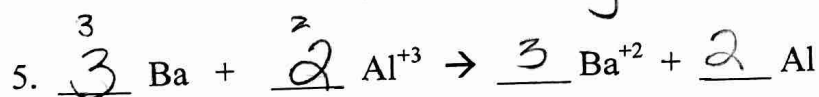
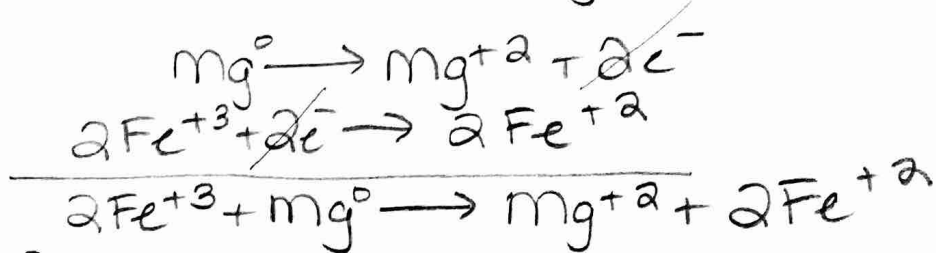
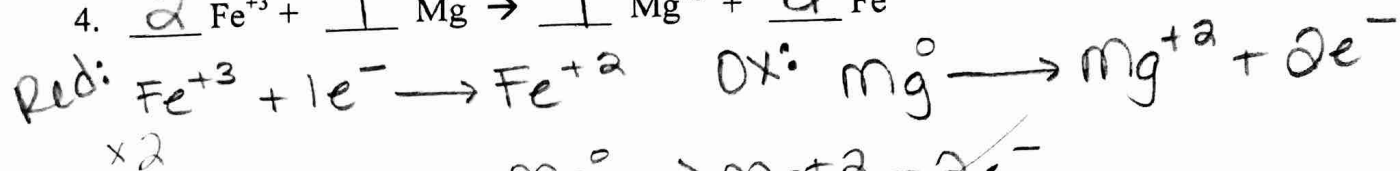
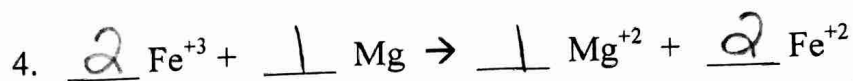


and



Part 2:

- ☐ Write balanced oxidation and reduction half reactions for each of the following (indicating which is which)
- ☐ Combine the half reactions in order to produce the balance the given net ionic equation.

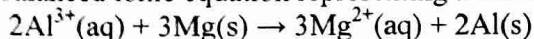


Redox and Half Reactions Regents Practice

Name: _____

1. Which balanced equation represents an oxidation-reduction reaction?
- (1) $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{NaCl}$ (3) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
- (2) $\text{C} + \text{H}_2\text{O} \rightarrow \text{CO} + \text{H}_2$ (4) $\text{Mg}(\text{OH})_2 + 2\text{HNO}_3 \rightarrow \text{Mg}(\text{NO}_3)_2 + 2\text{H}_2\text{O}$

2. Given the balanced ionic equation representing a reaction:

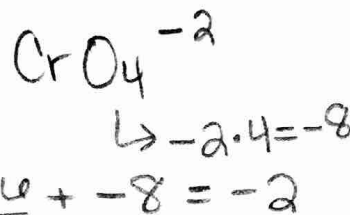


In this reaction, electrons are transferred from

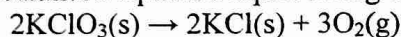
- (1) Al to Mg^{2+} (3) Mg to Al^{3+}
- (2) Al^{3+} to Mg (4) Mg^{2+} to Al

3. What is the oxidation number of chromium in the chromate ion, CrO_4^{2-} ?

- (1) +6 (3) +3
- (2) +2 (4) +8

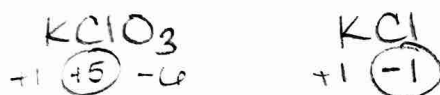


4. Given the balanced equation representing a reaction:

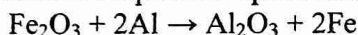


The oxidation state of chlorine in this reaction changes from

- (1) -1 to +1 (3) +1 to -1
- (2) -1 to +5 (4) +5 to -1

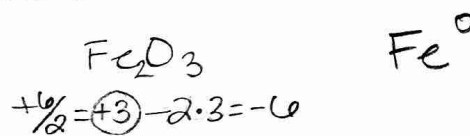


5. Given the balanced equation representing a reaction:



During this reaction, the oxidation number of Fe changes from

- (1) +2 to 0 as electrons are transferred (3) +3 to 0 as electrons are transferred
- (2) +2 to 0 as protons are transferred (4) +3 to 0 as protons are transferred



6. Which balanced equation represents a redox reaction?

- (1) $\text{PCl}_5 \rightarrow \text{PCl}_3 + \text{Cl}_2$ (3) $\text{LiBr} \rightarrow \text{Li}^+ + \text{Br}^-$
- (2) $\text{KOH} + \text{HCl} \rightarrow \text{KCl} + \text{H}_2\text{O}$ (4) $\text{Ca}^{2+} + \text{SO}_4^{2-} \rightarrow \text{CaSO}_4$

7. When lithium reacts with bromine to form the compound LiBr, each lithium atom

- (1) gains one electron and becomes a negatively charged ion
- (2) gains three electrons and becomes a negatively charged ion
- (3) loses one electron and becomes a positively charged ion
- (4) loses three electrons and becomes a positively charged ion

Base your answers to questions 15 and 16 on the information below.

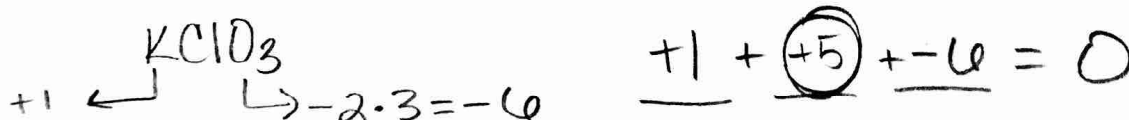
The unbalanced equation below represents the decomposition of potassium chlorate.



8. Balance the equation below, using the smallest whole-number coefficients.



9. Determine the oxidation number of chlorine in the reactant.



10. In an oxidation-reduction reaction, the number of electrons lost is
 (1) equal to the number of electrons gained
 (2) equal to the number of protons gained
 (3) less than the number of electrons gained
 (4) less than the number of protons gained

11. Given the balanced equation representing a reaction: $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + 55.8 \text{ kJ}$

In this reaction there is conservation of

- (1) mass, only
 (2) mass and charge, only
 (3) mass and energy, only
 (4) mass, charge, and energy

12. Given the unbalanced ionic equation: $3\text{Mg}^0 + \underline{2} \text{Fe}^{3+} \rightarrow 3\text{Mg}^{2+} + \underline{2} \text{Fe}^0$

When this equation is balanced, both Fe^{3+} and Fe have a coefficient of

- (1) 1, because a total of 6 electrons is transferred
 (2) 2, because a total of 6 electrons is transferred
 (3) 1, because a total of 3 electrons is transferred
 (4) 2, because a total of 3 electrons is transferred

13. Given the balanced equation representing a reaction: $\text{Mg}(\text{s}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Ni}(\text{s})$

What is the total number of moles of electrons lost by $\text{Mg}(\text{s})$ when 2.0 moles of electrons are gained by $\text{Ni}^{2+}(\text{aq})$?

- (1) 1.0 mol
 (2) 2.0 mol
 (3) 3.0 mol
 (4) 4.0 mol

14. Which half-reaction correctly represents reduction?

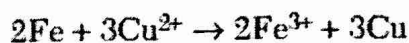
- (1) $\text{Mn}^{4+} \rightarrow \text{Mn}^{3+} + \text{e}^-$
 (2) $\text{Mn}^{4+} \rightarrow \text{Mn}^{7+} + 3\text{e}^-$
 (3) $\text{Mn}^{4+} + \text{e}^- \rightarrow \text{Mn}^{3+}$
 (4) $\text{Mn}^{4+} + 3\text{e}^- \rightarrow \text{Mn}^{7+}$

15. Which equation shows conservation of mass and charge?

- (1) $\text{NH}_4\text{Br} \rightarrow \text{NH}_3 + \text{Br}_2$
 (2) $2\text{Mg} + \text{Fe}^{3+} \rightarrow \text{Mg}^{2+} + 3\text{Fe}$
 (3) $\text{H}_2\text{SO}_4 + \text{LiOH} \rightarrow \text{Li}_2\text{SO}_4 + \text{H}_2\text{O}$
 (4) $\text{Cu} + 2\text{Ag}^+ \rightarrow \text{Cu}^{2+} + 2\text{Ag}$

16. Which half-reaction equation represents the reduction of a potassium ion?

- (1) $\text{K}^+ + \text{e}^- \rightarrow \text{K}$
 (2) $\text{K} + \text{e}^- \rightarrow \text{K}^+$
 (3) $\text{K}^+ \rightarrow \text{K} + \text{e}^-$
 (4) $\text{K} \rightarrow \text{K}^+ + \text{e}^-$



17. Given the balanced equation representing a reaction:

When the iron atoms lose six moles of electrons, how many moles of electrons are gained by the copper ions?

- (1) 12 moles
 (2) 2 moles
 (3) 3 moles
 (4) 6 moles

18. Which half-reaction equation represents the reduction of an iron(II) ion?

- (1) $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$
 (2) $\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$
 (3) $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$
 (4) $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$

Spontaneity of Redox Reactions: HOW DO YOU KNOW IF IT WILL HAPPEN?

Metals that are
up higher on
Table J are
more reactive

In order for rxn to
happen spontaneously
solid metal must be
up higher than
metal in solution

metal up higher on
chart is more likely
to be oxidized (become a + ion)

Table J
Activity Series**

Most Active	Metals	Nonmetals	Most Active
	Li	F ₂	
	Rb	Cl ₂	
	K	Br ₂	
	Cs	I ₂	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	H ₂		
	Cu		
	Ag		
	Au		
Least Active			Least Active

** Activity Series is based on the hydrogen standard. H₂ is not a metal.

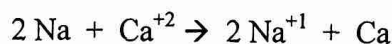
Spontaneity of Reactions- Table J Practice

Name: _____

1. Based on Table J, which of the following metals is most reactive?
 a. Ag b. Au ☒ c. Ca d. Cu

2. Based on Table J, which of the following metals is most likely to be oxidized?
☒ a. Li b. Na c. K d. Cs

3. Based on Table J, circle the reaction below that is going to be "spontaneous." Explain your choice:



Ca is more reactive/likely to be oxidized
 → it is up higher on Table J

4. Circle all of the pairs for which a spontaneous reaction will occur:

☒ a. Al, CuCl_2

☒ c. Sn, Pb

☒ e. Mg^{2+} , Co

☒ g. Li, CaO

~~b. Cu, HCl~~

~~d. Au, Ag^{+}~~

~~f. Ni^{2+} , Sn^{2+}~~

☒ h. H^{+} , Sn

5. Which metal is more active than Ni and less active than Zn?

~~a. Cu~~

~~b. Mg~~

☒ c. Cr

~~d. Pb~~

6. Which metal is more active than H_2 ?

a. Ag

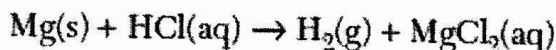
b. Cu

c. Au

☒ d. Pb

Use your answers to questions 7 through 9 on the information below.

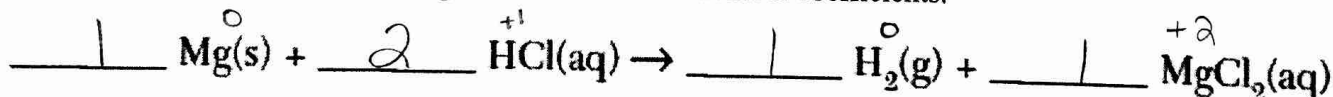
In a laboratory investigation, magnesium reacts with hydrochloric acid to produce hydrogen gas and magnesium chloride. This reaction is represented by the unbalanced equation below.



7. State, in terms of the relative activity of elements, why this reaction is spontaneous.

Magnesium is more reactive than Hydrogen
 so it will be oxidized → Mg is higher on Table J

8. Balance the equation below, using the smallest whole-number coefficients.



9. Write a balanced half-reaction equation for the oxidation that occurs.



10. On the non-metals side of Chart J, explain why it makes sense that F_2 is most reactive and I_2 is least. Explain using the definition of electronegativity, as well as the electronegativity values.

F_2 is more reactive because it is more electronegative and will fight harder to remove others' electrons.
 F electronegativity = 4.0 I electronegativity = 2.7 21