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)**

↳ electrical outlet

Electrolytic Cell

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

→ **Movable Charges from a CHEMICAL SOURCE (SOLUTION)**

↳ Ions in Sol'n

Voltaic Cell

↳ 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

1. What is the formula of titanium (II) oxide?
   a. TiO b. Ti_2O c. TiO_2 d. Ti_2O_3

2. What is the oxidation state of chlorine in NaCl?
   a. -1 b. +1 c. +3 d. +5

3. What is the oxidation state of nitrogen in $NaNO_2$?
   a. +1 b. +3 c. +2 d. +4

4. What is the oxidation number of chromium in the chromate ion, CrO_4^{-2}?
   a. +6 b. +3 c. +2 d. +8

5. Given the following reaction: $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$ 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

1. All chemical reactions have a conservation of
   a. mass, only
   b. charge and energy, only
   c. mass and charge, only
   d. mass, charge, and energy

2. Which equation shows conservation of both mass and charge?
   a. $Cl_2 + Br^- \rightarrow Cl^- + Br_2$
   b. $Zn + Cr^{3+} \rightarrow Zn^{2+} + Cr$
   c. $Cu + 2\,Ag^+ \rightarrow Cu^{2+} + Ag$
   d. $Ni + Pb^{2+} \rightarrow Ni^{2+} + Pb$

3. Given the balanced ionic equation: $2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$
   Compared to the total charge of the reactants, the total charge of the products is
   a. Less b. Greater c. The same

4. During which process does an atom gain one or more electrons?
   a. Transmutation
   b. Oxidation
   c. Reduction
   d. Neutralization

5. Which reaction is an example of an oxidation reduction reaction?
   a. $AgNO_3 + KI \rightarrow AgI + KNO_3$
   b. $Cu + 2\,AgNO_3 \rightarrow Cu(NO_3)_2 + 2\,Ag$
   c. $2\,KOH + H_2SO_4 \rightarrow K_2SO_4 + 2\,H_2O$
   d. $Ba(OH)_2 + 2\,HCl \rightarrow BaCl_2 + 2\,H_2O$

DR can NEVER be Redox!

Part 3- Identifying Species Oxidized and Reduced

1. When a neutral atom undergoes oxidation, the atom's oxidation state
   a. decreases as it gains electrons
   b. increases as it gains electrons
   c. decreases as it loses electrons
   d. increases as it loses electrons

2. Which change in oxidation number indicates oxidation?
   a. −1 to +2
   b. −1 to −2
   c. +2 to −3
   d. +3 to +2

3. Which changes occur when Pt^{2+} is reduced?
   a. The Pt^{2+} gains electrons and its oxidation number increases.
   b. The Pt^{2+} gains electrons and its oxidation number decreases.
   c. The Pt^{2+} loses electrons and its oxidation number increases.
   d. The Pt^{2+} loses electrons and its oxidation number decreases.

4. Given the equation: $C(s) + H_2O(g) \rightarrow CO(g) + H_2(g)$ Which species undergoes reduction?
   a. C(s)
   b. C^{2+}
   c. H^{+}
   d. $H_2(g)$

5. Given the equation: $3Au(s) + 2Fe^{+3}(aq) \rightarrow 3Au^{+2}(aq) + 2Fe(s)$
   a. Gold⁰ is reduced as it loses electrons
   b. Iron is reduced as it gains electrons
   c. Gold⁰ is oxidized as it loses electrons
   d. Iron is Oxidized as it gains electrons

(III) (III)

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: $Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq)$

Cell with zinc and iron electrodes: $Fe^{2+}(aq) + Zn(s) \rightarrow Fe(s) + Zn^{2+}(aq)$

1. 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.

   $Cu^{+2} \rightarrow Cu^{0}$ gained e^{-} / $Fe \rightarrow Fe^{+2}$

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

   electrons

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

   $Zn^{0} \rightarrow Zn^{+2}$ charge ↑

4. 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	GER
→ oxidation	→ reduction
→ lose e^-	→ gain e^-
→ charge ↑ (less neg)	→ charge ↓ (more neg)

Half-Reactions

$$\overset{+2}{Cu}\overset{-2}{SO_4}{}_{(aq)} + \overset{0}{Fe}_{(s)} \longrightarrow \overset{0}{Cu}_{(s)} + \overset{+2}{Fe}\overset{-2}{SO_4}{}_{(aq)}$$

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

Red: $Cu^{+2} + 2e^- \longrightarrow Cu^0$

Ox: $Fe^0 \longrightarrow Fe^{+2} + 2e^-$

NOT Balancing Half-Reactions

$$H_2^{0} \longrightarrow \underset{(1)}{\underline{\underline{2}}}H^{+1} + \underset{(2)}{\underline{\underline{2e^-}}}$$

(1) Balance for mass $\rightarrow$ coefficients

(2) Balance for charge $\rightarrow$ add e^-

NOTES

Balancing Net Ionic Equations

$$2\overset{+1}{Ag}\overset{-1}{Cl}_{(aq)} + \overset{0}{Zn}_{(s)} \longrightarrow \overset{+2}{Zn}\overset{-1}{Cl}_{2(aq)} + 2\overset{0}{Ag}_{(s)}$$

① Assign oxidation #'s

② Redox? who ox/red?

③ Write ½ rxns

ox: $Zn^0 \longrightarrow Zn^{+2} + 2e^-$

red: $Ag^{+1} + 1e^- \longrightarrow Ag^0$

④ Multiply ½ rxns so # e^- lost = # e^- gained

$2(Ag^{+1} + 1e^- \longrightarrow Ag^0)$ $\quad 1(Zn^0 \longrightarrow Zn^{+2} + 2e^-)$

$2Ag^{+1} + \underline{2e^-} \longrightarrow 2Ag^0$ $\quad Zn^0 \longrightarrow Zn^{+2} + \underline{2e^-}$

⑤ add ½ rxns to cancel out e^- and end w/ net ionic EQ

$$2Ag^{+1} + \cancel{2e^-} \longrightarrow 2Ag^0$$
$$+ \quad Zn^0 \longrightarrow Zn^{+2} + \cancel{2e^-}$$

$$\boxed{2Ag^{+1} + Zn^0 \longrightarrow 2Ag^0 + Zn^{+2}}$$ net ionic equation

Balancing Half-Reactions and Net Ionic Equations Practice Name:____________________

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**

1. Fe^{+3} + <u>1</u> e^{-1} → Fe^{+2} and Zn → Zn^{+2} + <u>2</u> e^{-1}

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

$2Fe^{+3} + 2e^{-} \rightarrow 2Fe^{+2}$ $Zn \rightarrow Zn^{+2} + 2e^{-1}$

$2Fe^{+3} + \cancel{2e^{-}} \rightarrow 2Fe^{+2}$

$+ \quad Zn \rightarrow Zn^{+2} + \cancel{2e^{-}}$

$\boxed{2Fe^{+3} + Zn \rightarrow Zn^{+2} + 2Fe^{+2}}$

2. Ca → Ca^{+2} + <u>2</u> e^{-1} and Al^{+3} + <u>3</u> e^{-1} → Al

$3(Ca \rightarrow Ca^{+2} + 2e^{-})$ $2(Al^{+3} + 3e^{-} \rightarrow Al)$

$3Ca \rightarrow 3Ca^{+2} + 6e^{-}$ $2Al^{+3} + 6e^{-} \rightarrow 2Al$

$3Ca \rightarrow 3Ca^{+2} + \cancel{6e^{-}}$

$+ \quad 2Al^{+3} + \cancel{6e^{-}} \rightarrow 2Al$

$\boxed{2Al^{+3} + 3Ca \rightarrow 2Al + 3Ca^{+2}}$

3. Sr^{+2} + $2e^{-}$ → Sr and Rb → Rb^{+1} + $1e^{-}$

$2(Rb \rightarrow Rb^{+1} + 1e^{-})$

$2Rb \rightarrow 2Rb^{+1} + 2e^{-}$

$Sr^{+2} + \cancel{2e^{-}} \rightarrow Sr$

$2Rb \rightarrow 2Rb^{+1} + \cancel{2e^{-}}$

$\boxed{Sr^{+2} + 2Rb \rightarrow Sr + 2Rb^{+1}}$

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.

4. $\underline{2}$ Fe^{+3} + $\underline{1}$ Mg → $\underline{1}$ Mg^{+2} + $\underline{2}$ Fe^{+2}

Red: $Fe^{+3} + 1e^- \rightarrow Fe^{+2}$ Ox: $Mg^0 \rightarrow Mg^{+2} + 2e^-$

×2

$$Mg^0 \rightarrow Mg^{+2} + \cancel{2e^-}$$
$$2Fe^{+3} + \cancel{2e^-} \rightarrow 2Fe^{+2}$$
$$2Fe^{+3} + Mg^0 \rightarrow Mg^{+2} + 2Fe^{+2}$$

5. $\underline{3}$ Ba + $\underline{2}$ Al^{+3} → $\underline{3}$ Ba^{+2} + $\underline{2}$ Al

Red: $Al^{+3} + 3e^- \rightarrow Al^0$ Ox: $Ba^0 \rightarrow Ba^{+2} + 2e^-$

×2 ×3

$$2Al^{+3} + \cancel{6e^-} \rightarrow 2Al^0$$
$$3Ba^0 \rightarrow 3Ba^{+2} + \cancel{6e^-}$$
$$2Al^{+3} + 3Ba^0 \rightarrow 3Ba^{+2} + 2Al^0$$

6. $\underline{4}$ Al + $\underline{3}$ Ti^{+4} → $\underline{4}$ Al^{+3} + $\underline{3}$ Ti

Red: $Ti^{+4} + 4e^- \rightarrow Ti^0$ Ox: $Al^0 \rightarrow Al^{+3} + 3e^-$

×3 ×4

$$3Ti^{+4} + \cancel{12e^-} \rightarrow 3Ti^0$$
$$4Al^0 \rightarrow 4Al^{+3} + \cancel{12e^-}$$
$$4Al + 3Ti^{+4} \rightarrow 4Al^{+3} + 3Ti$$

7. $\underline{2}$ H^{+1} + $\underline{1}$ Zn → $\underline{1}$ Zn^{+2} + $\underline{1}$ H_2

Red: $2H^{+1} + 2e^- \rightarrow H_2^0$ Ox: $Zn^0 \rightarrow Zn^{+2} + 2e^-$

$$2H^{+1} + \cancel{2e^-} \rightarrow H_2^0$$
$$Zn^0 \rightarrow Zn^{+2} + \cancel{2e^-}$$
$$2H^{+1} + Zn^0 \rightarrow Zn^{+2} + H_2^0$$

Redox and Half Reactions Regents Practice **Name:** ____________________

1. Which balanced equation represents an oxidation-reduction reaction?
   (1) $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + 2NaCl$
   (2) $C + H_2O \rightarrow CO + H_2$
   (3) $CaCO_3 \rightarrow CaO + CO_2$
   (4) $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$

2. Given the balanced ionic equation representing a reaction:
   $$2Al^{3+}(aq) + 3Mg(s) \rightarrow 3Mg^{2+}(aq) + 2Al(s)$$
   In this reaction, electrons are transferred from
   (1) Al to Mg^{2+}
   (2) Al^{3+} to Mg
   (3) Mg to Al^{3+}
   (4) Mg^{2+} to Al

3. What is the oxidation number of chromium in the chromate ion, CrO_4^{2-}?
   (1) +6
   (2) +2
   (3) +3
   (4) +8

   CrO_4^{-2}
   $\hookrightarrow -2 \cdot 4 = -8$
   $+6 + -8 = -2$

4. Given the balanced equation representing a reaction:
   $$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$
   The oxidation state of chlorine in this reaction changes from
   (1) −1 to +1
   (2) −1 to +5
   (3) +1 to −1
   (4) +5 to −1

   $KClO_3$: +1 +5 −6 KCl: +1 −1

5. Given the balanced equation representing a reaction:
   $$Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$$
   During this reaction, the oxidation number of Fe changes from
   (1) +2 to 0 as electrons are transferred
   (2) +2 to 0 as protons are transferred
   (3) +3 to 0 as electrons are transferred
   (4) +3 to 0 as protons are transferred

   Fe_2O_3 Fe^0
   $+6/2 = +3$ $-2 \cdot 3 = -6$

6. Which balanced equation represents a redox reaction?
   (1) $PCl_5 \rightarrow PCl_3 + Cl_2$
   (2) $KOH + HCl \rightarrow KCl + H_2O$
   (3) $LiBr \rightarrow Li^+ + Br^-$
   (4) $Ca^{2+} + SO_4^{2-} \rightarrow CaSO_4$

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.

$$KClO_3(s) \rightarrow KCl(s) + O_2(g)$$

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

   __2__ $KClO_3(s) \rightarrow$ __2__ $KCl(s)$ + __3__ $O_2(g)$

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

   $KClO_3$
   +1 $\hookrightarrow -2 \cdot 3 = -6$
   $+1 + (+5) + -6 = 0$

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: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l) + 55.8$ 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: $3Mg^0 + \underline{2}\ Fe^{3+} \rightarrow 3Mg^{2+} + \underline{2}\ Fe^0$

+6 | +6

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: $Mg(s) + Ni^{2+}(aq) \rightarrow Mg^{2+}(aq) + Ni(s)$

What is the total number of moles of electrons lost by Mg(s) when 2.0 moles of electrons are gained by $Ni^{2+}(aq)$?
(1) 1.0 mol
(2) 2.0 mol
(3) 3.0 mol
(4) 4.0 mol

14. Which half-reaction correctly represents reduction?
(1) $Mn^{4+} \rightarrow Mn^{3+} + e^-$
(2) $Mn^{4+} \rightarrow Mn^{7+} + 3e^-$
(3) $Mn^{4+} + e^- \rightarrow Mn^{3+}$
(4) $Mn^{4+} + 3e^- \rightarrow Mn^{7+}$

charge ↓
gain e⁻

15. Which equation shows conservation of mass and charge?
(1) $NH_4Br \rightarrow NH_3 + Br_2$
(2) $2Mg + Fe^{3+} \rightarrow Mg^{2+} + 3Fe$
(3) $H_2SO_4 + LiOH \rightarrow Li_2SO_4 + H_2O$
(4) $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$

16. Which half-reaction equation represents the reduction of a potassium ion?
(1) $K^+ + e^- \rightarrow K$
(2) $K + e^- \rightarrow K^+$
(3) $K^+ \rightarrow K + e^-$
(4) $K \rightarrow K^+ + e^-$

17. Given the balanced equation representing a reaction: $2Fe + 3Cu^{2+} \rightarrow 2Fe^{3+} + 3Cu$
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?

+2

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

charge ↓
gain e⁻

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

Table J
Activity Series**

Most Active ↓ Least Active	Metals	Nonmetals	Most Active ↓ Least 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		

**Activity Series is based on the hydrogen standard. H_2 is not a metal.

Metals that are up higher on Table J are more reactive

- Inorder 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)

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:

   $Ca + 2\,Na^{+1} \rightarrow Ca^{+2} + 2\,Na$ OR $2\,Na + Ca^{+2} \rightarrow 2\,Na^{+1} + Ca$

   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

ase 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.

$$Mg(s) + HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$$

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.

   $\underline{\;1\;}\ \overset{0}{Mg}(s) + \underline{\;2\;}\ \overset{+1}{H}Cl(aq) \rightarrow \underline{\;1\;}\ \overset{0}{H_2}(g) + \underline{\;1\;}\ \overset{+2}{Mg}Cl_2(aq)$

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

   $Mg^0 \longrightarrow Mg^{+2} + 2e^-$

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 electrone[g]
    and will fight harder to remove others' electrons (
    F electroneg. = 4.0 I electroneg = 2.7