Laney College Chem 30A
Professor Scott Beaver

Homework #6
3/17/2016
Name: Scott Beaver

1. Balance the following reactions:

a. $\_2\_\ Fe\ (s) + \_\frac{3}{2}\_\ O_2\ (g) \rightarrow \_\_\ Fe_2O_3\ (s)$ (2x) (2x)

$4\ Fe + 3O_2 \rightarrow 2\ Fe_2O_3$

b. $\_\_\_\ H_2SO_4\ (aq) + \_2\_\ LiOH\ (aq) \rightarrow \_\_\ Li_2SO_4\ (aq) + \_2\_\ H_2O\ (l)$

c. $\_2\_ NH_4I\ (aq) + \_\_ Pb(NO_3)_2\ (aq) \rightarrow \_\_\ PbI_2\ (s) + \_2\_\ NH_4NO_3\ (aq)$

Pb^{2+} NO_3^{1-} ✗ error – it's $Pb(NO_3)_2$ 2

d. $\_\_\_\ C_2H_2\ (g) + \_\frac{5}{2}\_\ O_2\ (g) \rightarrow \_2\_\ CO_2\ (g) + \_1\_\ H_2O\ (g)$ (2x) (2x) 5 O

"odd oxygen" problem

$2\ C_2H_2 + 5O_2 \rightarrow 4\ CO_2 + 2\ H_2O$

10 O

e. $\_\_\_\ C_3H_8\ (g) + \_5\_\ O_2\ (g) \rightarrow \_3\_\ CO_2\ (g) + \_4\_\ H_2O\ (g)$

an "even oxygen" problem

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2. Determine if precipitation occurs for each equation below. Use the table of solubility rules on the back page of this homework. (Please remove the table for use, and do not submit a printed copy back to me. I have plenty of solubility tables already, thanks.)

If a precipitation reaction occurs:

1. Complete the products with physical states (*s, l, aq, g*),
2. balance the equation, and
3. provide the net ionic equation below the balanced equation.

If no precipitant (insoluble chemical species) forms, the correct answer is simply "n.r." meaning that no reaction occurs.

$AgNO_3\ (aq) + HCl\ (aq) \rightarrow AgCl(s) + HNO_3(aq)$

$AgCl$ - s
HNO_3 - aq (acid)

Net ionic: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

$AgNO_3\ (aq) + NaBr\ (aq) \rightarrow AgBr(s) + NaNO_3(aq)$

$AgBr$ - s
$NaNO_3$ - aq

Net ionic: $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$

$AgNO_3\ (aq) + Na_2SO_4\ (aq) \rightarrow$ n.r.

Ag_2SO_4 - aq
$NaNO_3$ - aq

$Ba(NO_3)_2\ (aq) + Na_2SO_4\ (aq) \rightarrow BaSO_4(s) + 2\,NaNO_3(aq)$

$BaSO_4$ - s
$NaNO_3$ - aq

Net ionic: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

$NH_4Cl\ (aq) + KCH_3COO\ (aq) \rightarrow$ n.r.

NH_4CH_3COO - aq
KCl - aq

$2\,NaCl\ (aq) + PbSO_4\ (aq) \rightarrow PbCl_2(s) + Na_2SO_4(aq)$

$PbCl_2$ - s
Na_2SO_4 - aq

Net ionic: $Pb^{2+}(aq) + 2\,Cl^-(aq) \rightarrow PbCl_2(s)$

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3. Complete and balance the neutralization reactions. Write out the net ionic reaction for each in the space provided.

a. ___ $HCl\ (aq)$ + ___ $NaOH\ (aq)$ → $H_2O\ (l)$ + $NaCl\ (aq)$

Net ionic equation:

$H^+ + OH^- \rightarrow H_2O$

b. ___ $H_2SO_4\ (aq)$ + 2 $NaOH\ (aq)$ → $2H_2O\ (l)$ + $Na_2SO_4\ (aq)$

Net ionic equation: $2H^+ + 2\ OH^- \rightarrow 2H_2O$

c. ___ $H_3PO_4\ (aq)$ + 3 $KOH\ (aq)$ → $3H_2O\ (l)$ + $K_3PO_4\ (l)$

Net ionic equation: $3H^+ + 3OH^- \rightarrow 3H_2O$

d. 2 $HNO_3\ (aq)$ + ___ $Mg(OH)_2\ (aq)$ → $2H_2O_{(l)}$ + $Mg(NO_3)_2\ (aq)$

Net ionic equation: $2H^+ + 2OH^- \rightarrow 2H_2O$

e. ___ $HCH_3COO\ (aq)$ + ___ $NaOH\ (aq)$ → $H_2O_{(l)}$ + $NaCH_3COO\ (aq)$

≡ the acid is H^+

Net ionic equation:

$H^+ + OH^- \rightarrow H_2O$

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O_2 is most like oxygen, hence it's the oxidizing agent (more electronegative)

4. Circle the chemical species that is:

a. the oxidizing agent (reduced – oxid # ↓)

$2\,Mg\,(s) + O_2\,(g) \rightarrow 2\,MgO\,(s)$ — O goes from 0 to −2

$Zn\,(s) + Cu^{2+}\,(aq) \rightarrow Zn^{2+}\,(aq) + Cu\,(s)$ — Cu goes from +2 → 0

$2\,NO\,(g) + 2\,CO\,(g) \rightarrow N_2\,(g) + 2\,CO_2\,(g)$ — N goes from +2 → 0

for CO → CO_2, the carbonaceous species gains O atoms. It is oxidized

b. reduced (again, oxid # ↓. Same as the oxidizing agent.)

$2\,C\,(s) + O_2\,(g) \rightarrow 2\,CO\,(g)$ — O goes 0 to −2 (like oxygen)

$H_2O\,(s) + CO\,(g) \rightarrow H_2\,(g) + CO_2\,(g)$ — H goes +1 to 0

$I^-\,(aq) + ClO^-\,(aq) \rightarrow IO^-\,(aq) + Cl^-\,(aq)$ — Cl goes +1 to −1

$2\,NaClO_2\,(s) + Cl_2\,(g) \rightarrow 2\,ClO_2\,(g) + 2\,NaCl\,(s)$ — tricky — BONUS material

ClO_2^- (aq), as sodium hypochlorite, is bleach, a common oxidizing agent.

chlorine has positive oxidation # because O (EN = 3.5) is more electronegative than Cl (EN = 3.0). Typically, Cl would be −1 unless w/ a more electronegative atom.

5. Calculate the oxidation state of carbon in the following:

CO_2	+4	CH_2O	0
CH_4	−4	HCN	+2
C_2H_6	−3	CO_3^{2-}	+4
CCl_4	+4	CBr_4	+4

$\frac{-6}{2} = -3$

CO_2 and CCl_4 have same oxidation state for C, because it forms 4 bonds w/ more electronegative atoms. CBr_4 same as CCl_4.