**Unit 11: Acids & Bases and Radiochemistry**

**Ch. Acids, Bases, and Radioactivity**

<table>
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<th>Common Strong Acids</th>
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<td>HCl</td>
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<td>HNO₃</td>
<td>nitric acid</td>
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<tr>
<td>H₂SO₄</td>
<td>sulfuric acid</td>
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<tr>
<td>HClO₄</td>
<td>perchloric acid</td>
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<tr>
<td>If #O - #H ≥ 2, then strong acid.</td>
<td>Sr(OH)₂</td>
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<table>
<thead>
<tr>
<th>Common Weak Acids</th>
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<td>HC₂H₃O₂ = CH₃CO₂H</td>
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<td>** H₂NOH hydroxyl amine</td>
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<td>** H₂NNH₂ hydrazine</td>
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</table>

**Acids: start with “H” or end in –CO₂H (-COOH) and have a sour taste.**

**Bases: end in –OH (positive ion in front) or end in –NH, -NH₂, NH₃**

They have a bitter taste and feel slippery.

* 3 Acid/Base Theories

I. **Arrhenius Theory**

a) **Acids produce H⁺ in a water solution**

$$\text{H₂O + HCl}(g) \rightarrow \text{H⁺(aq)} + \text{Cl⁻(aq)} (+ \text{H₂O})$$

b) **Bases produce OH⁻ in a water solution**

$$\text{H₂O + NaOH(s)} \rightarrow \text{Na⁺(aq)} + \text{OH⁻(aq)} (+ \text{H₂O})$$

Acids, Bases and Salts are Electrolytes (Ion forming in water).
II. Brønsted-Lowry Theory

a) Acids donate protons (H⁺) ← lose 1e⁻, 1 p left, 0 neutrons

\[
\text{HCl}(g) + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)
\]

\[\text{H}_3\text{O}^+ = \text{hydronium ion}\]

\[
\text{conjugate acid} \quad \text{conjugate base}
\]

(acid always has “H⁺")    - deprotonated acid

-biprotonated base

b) Bases accept protons

\[
\text{NH}_3(g) + \text{H}_2\text{O} \rightarrow \text{NH}_4^+(aq) + \text{OH}^-(aq)
\]

\[\text{conj. acid} \quad \text{conj. base} \quad \text{(has “H⁺")}\]

III. Lewis Theory

a) Acids are e⁻ pair acceptors.

b) Bases are e⁻ pair donors.

\内饰_\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{H:NH}_3^+ + \text{OH}^- \]

* Bases have extra electrons.

* Strong acids/bases completely ionize (fall apart).

* Weak acids/bases only partially ionize (equilibriums).

\[
\text{HF}(g) + \text{H}_2\text{O}(l) \leftrightarrow \text{H}_3\text{O}^+(aq) + \text{F}^-(aq) \quad \text{K}_a = 7.2 \times 10^{-4}
\]

\[
\text{K}_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}
\]

The larger the Kₐ, the more H₃O⁺, the stronger the acid.

(The larger the Kₐ, the more OH⁻, the stronger the base.)

Ex. 1) List the following in order of increasing acid strength: H₂O, HCl, & HF.

**Kₜₐₜ = 1 \times 10^{-14}

\[
\text{H}_2\text{O} (\text{K} = 1 \times 10^{-14}) < \text{HF} (\text{K} = 7.2 \times 10^{-4}) < \text{HCl (no K, so strong acid)}
\]

smallest K, so weakest

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IV. pH scale
- is a measure of the acidity (H₃O⁺ or H⁺)

← 0 − 7 − 14 →
acid     base

* pH = -log[H₃O⁺]     * pOH = -log[OH⁻]

** H₃O⁺ = H⁺

H₂O(l) + H₂O(l) ⇌ H₃O⁺(aq) + OH⁻(aq)

* K_w = 1 x 10⁻¹⁴ = [H₃O⁺][OH⁻]

-pH = pOH + pOH

Ex. 1) Find [H₃O⁺], [OH⁻], pH, & pOH for a strong acid:

a) 0.213 M HNO₃ (no K so strong acid)

\[
\begin{array}{ccc}
\text{HNO}_3 + \text{H}_2\text{O} & \rightarrow & \text{H}_3\text{O}^+ + \text{NO}_3^- \\
0.213 \text{ M} & 0 & 0 \\
-0.213 \text{ M} & +0.213 \text{ M} & +0.213 \text{ M} \\
0 & 0.213 \text{ M} & 0.213 \text{ M} \\
\end{array}
\] (Initial)

H₃O⁺ = 0.213 M

pH = -log(H₃O⁺) = -log(0.213 M) = -(-0.672) = 0.672

K_w = [H₃O⁺][OH⁻]
1 x 10⁻¹⁴ = (0.213 M)(OH⁻)

4.69 x 10⁻¹⁴ M = [OH⁻]  
pOH = -log(OH⁻) = -log(4.69 x 10⁻¹⁴ M)  
= -(-13.329) = 13.329

**The whole number is the power and is not significant!

OR

From pH = 0.672  
0.672 + pOH = 14

pOH = 13.328

4.70 x 10⁻¹⁴ = [OH⁻]
Ex. 2) Calculate $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ from:

a) $\text{pH} = 4.90$  \hspace{1cm} ← $\text{pH} < 7$ so acidic

$pH = - \log [\text{H}_3\text{O}^+]$

$4.90 = - \log [\text{H}_3\text{O}^+]$

$-4.90 = \log [\text{H}_3\text{O}^+]$  **move (-) before 10x

$10^x (-4.90 = \log [\text{H}_3\text{O}^+])$

$1.3 \times 10^{-5}$ M = $[\text{H}_3\text{O}^+]$

$K_w = 1 \times 10^{-14} = (1.3 \times 10^{-5}$ M) (OH$^-$)

$7.7 \times 10^{-10}$ M = $[\text{OH}^-]$

b) $\text{pOH} = 5.80$

$pOH = - \log [\text{OH}^-]$

$5.80 = - \log [\text{OH}^-]$

$10^x (-5.80 = \log [\text{OH}^-])$

$1.6 \times 10^{-6}$ M = $[\text{OH}^-]$

$K_w = 1 \times 10^{-14} = (\text{H}_3\text{O}^+) (1.6 \times 10^{-6}$ M)

$6.3 \times 10^{-9}$ M = $[\text{H}_3\text{O}^+]$

Is this acidic or basic? $\text{pH} = - \log [\text{H}_3\text{O}^+] = - \log 6.3 \times 10^{-9}$ M = **8.20** (pH > 7 so, basic)

Or look at $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$:

if $[\text{H}_3\text{O}^+]$ is larger, it is acidic,

if $[\text{OH}^-]$ is larger, it is basic.

** $\text{H}_3\text{O}^+ = \text{H}^+$**
V. Weak Acids/ Bases
-only partially ionize (equilibrium)

Ex. 1) What is the pH of a 0.200 M solution of acetic acid (\( \text{CH}_3\text{COOH} = \text{HC}_2\text{H}_3\text{O}_2^- \))?

(weak acid) What are the major species present?

\[
\begin{align*}
\text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O}_\text{(l)} & \leftrightarrow \text{H}_3\text{O}^+ + \text{C}_2\text{H}_3\text{O}_2^- \\
0.200 \text{ M} & \quad \text{(Initial)} \\
- x & \quad + x \quad + x \quad \text{(Change)} \\
0.200 \text{ M} - x & \quad x \quad x \quad \text{(Equilibrium)}
\end{align*}
\]

\[
\text{K}_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} 
\] no liquids: \( \text{H}_2\text{O}_\text{(l)} \)

\[
1.8 \times 10^{-5} = \frac{(x)(x)}{0.200 \text{ M} - x} \quad \text{If } \text{K}_a \text{ is smaller than the Concentration by } 10^3 \text{ or more, assume } x \text{ is small. (0.200 M} - x \text{ would equal just 0.200 M)}
\]

\[
1.8 \times 10^{-5} = \frac{x^2}{0.200 \text{ M}}
\]

\[
1.9 \times 10^{-3} \text{ M} = x = [\text{H}_3\text{O}^+] = [\text{C}_2\text{H}_3\text{O}_2^-]
\]

\[
\text{pH} = -\log(1.9 \times 10^{-3}) = 2.72
\]

Ex. 2) What is the pH of a 0.520 M solution of ethyl amine, \( \text{C}_2\text{H}_5\text{NH}_2 \)?

(weak base)

\[
\begin{align*}
\text{C}_2\text{H}_5\text{NH}_2 + \text{H}_2\text{O} & \leftrightarrow \text{C}_2\text{H}_5\text{NH}_3^+ + \text{OH}^- \\
0.520 \text{ M} & \quad \text{(Initial)} \\
-x & \quad +x \quad +x \quad \text{(Change)} \\
0.520 -x & \quad x \quad x \quad \text{(Equilibrium)}
\end{align*}
\]

\[
\text{K}_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]} 
\]

\[
5.6 \times 10^{-4} = \frac{(x)(x)}{(0.520 -x)} \quad x \text{ is small}
\]

\[
5.6 \times 10^{-4} = \frac{x^2}{(0.520)}
\]

\[
1.71 \times 10^{-2} = x = \text{OH}^- 
\]

\[
p\text{OH} = -\log[\text{OH}^-] 
\]

\[
\text{pOH} = -1.77 \quad 14 = \text{pH} + \text{pOH} 
\]

\[
14 - 1.77 = 12.23 = \text{pH}
\]
Ex. 3) What is the pH of a 0.340 M solution of HClO₂ (chlorous acid)

\[ K_{\text{HClO}_2} = 1.2 \times 10^{-2} \]

\[
\text{HClO}_2 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{ClO}_2^- \\
0.340 \text{ M} \\
-x \quad +x \quad +x \\
0.340 - x \quad x \quad x
\]

Kₐ = \[
\frac{[\text{H}_3\text{O}^+][\text{ClO}_2^-]}{[\text{HClO}_2]} \\
1.2 \times 10^{-2} = \frac{(x)(x)}{(0.340 - x)(3.40 \times 10^{-1})} \\
\text{x is not small}
\]

0.012 (0.340 - x) = x²

0.00408 - 0.012x - x² = 0

a = -1
b = -0.012
\[ c = 0.00408 \]

\[ \frac{-(-0.012) \pm \sqrt{(-0.012)^2 - 4(-1)(0.00408)}}{2(-1)} \]

\[ 0.012 \pm \sqrt{0.000144 + 0.01632} \]

\[ 0.012 \pm \sqrt{0.016464} \]

\[ 0.012 \pm 0.1283 = 0.1403 \text{ or } -0.1163 \]

x = -0.070 or 0.058

negative concentrations are impossible!

\[ [\text{H}_3\text{O}^+] = x = 0.058 \text{ M}, \text{ pH} = -\log [\text{H}_3\text{O}^+] = 1.24 \]

check: 
\[ K_a = (0.058)(0.058) = 1.19 \times 10^{-2} \approx 1.2 \times 10^{-2} \]

\[ \frac{(0.340 - 0.058)}{0.340 - 0.058} \]
Ex. 4) Find pH of 0.10 M Sr(OH)\textsubscript{2}.

**Strong Bases, like strong acids completely fall apart.**

\[
\text{Strong Base: } \text{Sr(OH)}_2 + \text{H}_2\text{O} \rightarrow \text{Sr}^{2+\text{(aq)}} + 2 \text{OH}^-\text {(aq)}
\]

\[
\begin{array}{ccc}
0.10 \text{ M} & - & 0.10 \text{ M} \\
0 & + & 0.10 \text{ M}
\end{array}
\]

\[
0 \quad 0.10 \text{ M} \quad 0.20 \text{ M}
\]

\[
[\text{OH}^-] = 0.20 \text{ M} \quad \text{pOH} = - \log [\text{OH}^-] = 0.70 \quad 14 - \text{pOH} = \text{pH} = 13.30
\]
#92 Notes  Ch. Aqueous Equilibria

I. Buffered Solutions

- resist changes in pH.

A buffered solution may contain a weak acid with its salt (CB) or a weak base with its salt (CA).

**Acids:**  
\[ \text{HNO}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NO}_2^- \]

\[ \text{CA} \quad \text{CB} \quad + (\text{Na}^+ \text{ or K}^+) \]

NaNO\(_2\) or KNO\(_2\) = salts of HNO\(_2\)

**Bases:**  
\[ \text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+ + \text{OH}^- \]

\[ \text{CA} + (\text{Cl}^- \text{ or Br}^-) \quad \text{CB} \]

C\(_6\)H\(_5\)NH\(_3\)Cl or C\(_6\)H\(_5\)NH\(_3\)Br = salts of C\(_6\)H\(_5\)NH\(_2\)

**Example 1:** What is the pH of a 0.300 M solution of HClO buffered with 0.0400 M NaClO?

K of HClO = 3.5 \times 10^{-8}

\[ \text{HClO} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{ClO}^- \]

0.300 M 0.0400 M ← from 0.0400 M Na\(^+\)ClO\(^-\)

-\(x\)  +\(x\)  +\(x\)

0.300 \(-x\) \(x\) 0.0400 \(+x\)

\[ 3.5 \times 10^{-8} = \frac{(x)(0.0400 + x)}{(0.300 - x)} \quad \text{x is small} \]

\[ 3.5 \times 10^{-8} = \frac{(x)(0.0400)}{(0.300)} \]

\[ 2.6 \times 10^{-7} = x = [\text{H}_3\text{O}^+] \quad \text{pH} = -\log [\text{H}_3\text{O}^+] = 6.58 \]
Ex. 2) What is the pH of a solution containing 0.450 M $H_2NNH_2$ and 0.250 M $H_2NNH_3Cl$. K of $H_2NNH_2 = 3.0 \times 10^{-6}$

$H_2NNH_2 + H_2O \leftrightarrow H_2NNH_3^+ + OH^-$

0.450 M 0.250 M ← from 0.250 M $H_2NNH_3^+Cl^-$

-x +x +x

0.450 – x 0.250 + x x

$3.0 \times 10^{-6} = \frac{(0.250 + x)(x)}{(0.450 \text{ - x})}$ x is small

$3.0 \times 10^{-6} = \frac{(0.250)(x)}{(0.450)}$

$5.4 \times 10^{-6} = x = [OH^-]$

pOH = $-\log [OH^-] = 5.27$

$pH = 14 - \text{pOH} = 8.73$
#93 Notes  II. Solubility Equilibria
- describes the amount of solid that dissolves in a saturated solution
  (no more solid will dissolve).

Ex. 1) Calculate the solubility of Ag₂CO₃ in mols per liter.  \( K_{sp} = 8.1 \times 10^{-12} \)

\[
\text{Ag}_2\text{CO}_3(s) \leftrightarrow 2 \text{Ag}^{+}(aq) + \text{CO}_3^{2-}(aq)
\]

\[
-x \quad +2x \quad +x
\]

\[
K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}]^1
\]

\[
K_{sp} = (2x)^2 (x)^1
\]

\[
K_{sp} = 4x^3
\]

\[
8.1 \times 10^{-12} = 4x^3
\]

\[
1.3 \times 10^{-4} \text{M} = x \quad \text{1.3} \times 10^{-4} \text{ mol/L}
\]

Ex. 2) Calculate the \( K_{sp} \) of CaF₂, if its solubility is \( 2.15 \times 10^{-4} \text{ mol/L} \).

\[
\text{CaF}_2(s) \leftrightarrow \text{Ca}^{2+}(aq) + 2 \text{F}^-(aq)
\]

\[
-x \quad x \quad 2x \quad x = 2.15 \times 10^{-4} \text{ mol/L}
\]

\[
K_{sp} = [\text{Ca}^{2+}]^1 [\text{F}^-]^2
\]

\[
K_{sp} = (x)^1 (2x)^2
\]

\[
K_{sp} = 4x^3
\]

\[
K_{sp} = 4 \times (2.15 \times 10^{-4} \text{ mol/L})^3 = 3.97 \times 10^{-11}
\]

Ex. 3) Calculate the solubility of Cu₃(PO₄)₂ in mol/L.  \( K_{sp} = 8.4 \times 10^{-15} \)

\[
\text{Cu}_3\text{(PO}_4\text{)}_2(s) \leftrightarrow 3 \text{Cu}^{2+}(aq) + 2 \text{PO}_4^{3-}(aq)
\]

\[
-x \quad +3x \quad +2x
\]

\[
K_{sp} = [\text{Cu}^{2+}]^3 [\text{PO}_4^{3-}]^2
\]

\[
K_{sp} = (3x)^3 (2x)^2
\]

\[
K_{sp} = (27 x^3) (4x^2)
\]

\[
K_{sp} = 108 x^5
\]

\[
8.4 \times 10^{-15} = 108 x^5
\]

\[
7.8 \times 10^{-17} = x^5
\]

\[
6.0 \times 10^{-4} \text{ mol/L} = x
\]
#94 Notes III. Precipitation Reactions

Ex.1) Will a precipitate form, if 25 ml of 0.100 M AgNO₃ is added to 105 ml of 0.200 M Na₂CO₃?

\[
\text{AgNO}_3(aq) + \text{Na}_2\text{CO}_3(aq) \rightarrow \text{Ag}_2\text{CO}_3(s) + \text{NaNO}_3
\]

\[\begin{align*}
\text{Ag}^{+} & \quad \text{NO}_3^{-} & \quad \text{Na}^{+} & \quad \text{CO}_3^{2-} \\
\quad & \quad \uparrow & \quad \rightarrow & \quad \uparrow
\end{align*}\]

\[K_{sp} = 8.1 \times 10^{-12} \quad \text{no } K_{sp}\]

a) Find concentration of ions in new volume:

\[\begin{align*}
\text{ML} & = \text{ML} \\
(0.100 \text{ M AgNO}_3) (25 \text{ ml}) & = M (130 \text{ ml}) \\
0.0192 \text{ mol/L} & = M \text{AgNO}_3 \\
\text{X1Ag}^+ & \quad \text{X1 NO}_3^- \\
0.0192 \text{ M Ag}^+ & \quad 0.0192 \text{ M NO}_3^- \\
\text{(0.200 M Na}_2\text{CO}_3) (105 \text{ ml}) & = M (130 \text{ ml}) \\
0.162 \text{ mol/L} & = M \text{Na}_2\text{CO}_3 \\
\text{X2Na}^+ & \quad \text{X1CO}_3^{2-} \\
0.324 \text{ M Na}^+ & \quad 0.162 \text{ M CO}_3^{2-}
\end{align*}\]

b) Put concentrations into \( K_{sp} \) equation:

\[\text{Ag}_2\text{CO}_3(s) \leftrightarrow 2 \text{Ag}^+ (aq) + \text{CO}_3^{2-} (aq)\]

\[K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}] = (0.0192 \text{ M Ag}^+)^2 (0.162 \text{ M CO}_3^{2-}) = 5.97 \times 10^{-5}\]

\[\text{K}_{sp \text{ calculated}} = 5.97 \times 10^{-5} \quad \text{>> Real } K_{sp \text{ from book}} = 8.1 \times 10^{-12}\]

since \( K_{sp} \) is more than saturated, solid forms (precipitate)

If the \( K_{sp} \) calculated is less than the real \( K_{sp} \),
it is less than saturated (not full), so no solid forms.
The “attractive” nuclear force between protons and neutrons is usually stronger than the repulsion energy of the protons. But some isotopes have too many protons and not enough neutrons, so there is too much repulsion. These atoms are unstable (radioactive). Atoms with more than 83 protons are always unstable (atomic #'s > 83).

I. Nuclear Decay Reactions (to become stable)

alpha (α) particle $^2_4$He - barely passes through paper

beta (β) particle $^-1_0e$ - stopped by 3 mm Al or 10 mm wood

gamma (γ) is electromagnetic energy - stopped by 60 cm Al or 7 cm Pb

positron particle $^+1_0e$

emission (given off), so put particle on product side

capture (taken in), so put particle on the reactant side

Ex. 1) $^{265}$Bh decays by α-emission.

$^{107}^{265}$Bh $\rightarrow$ $^2_4$He + $^{105}_{83}$Db

Ex. 2) $^{99}$Mo decays by β-emission (β-).

$^{42}_{99}$Mo $\rightarrow$ $^-1_0e + ^{43}_{83}$Tc

Ex. 3) $^{207}$Bi decays by positron emission (β+).

$^{83}_{207}$Bi $\rightarrow$ $^+1_0e + ^{82}_{207}$Pb

Ex. 4) $^{102}$Rh decays by electron capture (e.c. or ε).

$^{45}_{102}$Rh + $^-1_0e$ $\rightarrow$ $^{44}_{45}$Ru

II. Bombardments

-linear or cyclotron particle accelerators can be used to combine smaller atoms into larger atoms or to break up large atoms into smaller atoms.

Ex. 1) $^{249}$Bk is bombarded by $^{22}$Ne to make a new heavier element and 4 neutrons ($^4_0n$).

Write the reaction.

$^{97}_{249}$Bk + $^{10}_{22}$Ne $\rightarrow$ $^4_0n + ^{107}_{83}$Db
The rate of radioactive decay reactions is 1st order, so

\[
\text{Rate} = k [A]^1 \quad \text{←decays/sec}
\]

\[
t_{1/2} = \frac{0.693}{k} \quad \{t_{1/2} \text{ is the half-life (not } t \div 2)\}
\]

\[
\ln [A] = -kt + \ln [A]_0 \quad \{\text{order in textbook equation may be } \ln[A]_0 = kt + \ln[A]\}
\]

↑

↑

amount left after initial

time = t amount

Ex. 1) 28 g of $^{87}\text{Sr}$ has a 1/2 life of 2.8 hrs.

a) Find the decays/sec (decays = atoms).

\[
\begin{align*}
\text{Rate} &= k [A]^1 \\
t_{1/2} &= 0.693 / k \\
2.8 \text{ hr} &= 0.693 / k \\
10080 \text{ sec} &= 0.693 / k \\
(10080) k &= 0.693 \\
k &= 6.875 \times 10^{-5} \text{ sec}^{-1}
\end{align*}
\]

\[
\text{Rate} = k [A]^1 = (6.875 \times 10^{-5} \text{ sec}^{-1}) (28 \text{ g}^{87}\text{Sr} \frac{1 \text{ mol}}{1 \text{ mol}} \frac{6.022 \times 10^{23} \text{ atoms}}{87 \text{ g}})
\]

\[
\text{Rate} = 1.3 \times 10^{19} \text{ decays/sec}
\]

b) How much is left after 4.0 hrs?

\[
\begin{align*}
t_{1/2} &= 2.8 \text{ hr} \\
t_{1/2} &= 0.693 / k \\
2.8 \text{ hr} &= 0.693 / k \\
(2.8) k &= 0.693 \\
k &= 0.2475 \text{ hr}^{-1}
\end{align*}
\]

\[
\ln A = -kt + \ln A_0
\]

\[
\ln A = -(0.2475 \text{ hr}^{-1})(4.0 \text{ hr}) + \ln (28 \text{ g})
\]

\[
\ln A = -0.99 + 3.33
\]

\[
e^x (\ln A = 2.34)
\]

\[
A = 10. \text{ g is left}
\]
c) How much decayed in 4.0 hrs?

28 g \textit{initially} – 10. g \textit{left} = \textbf{18 g decayed}

Ex.2) How long does it take for 62.3\% of \(^{133}\text{Xe}\) to decay? (1/2 life = 5.3 days)

\[ t_{\frac{1}{2}} = \frac{0.693}{k} \]

5.3 d = 0.693/k

(5.3 d) k = 0.693

\[ k = 0.131 \text{ d}^{-1} \quad 100\% \text{ initially} \quad = [A]_o \]

\[ \frac{-62.3\% \text{ decays}}{37.7\% \text{ is left}} \quad = [A] \]

\[ \ln [A] = -kt + \ln [A]_o \]

\[ \ln (37.7) = -(0.131 \text{ d}^{-1}) t + \ln (100) \]

3.63 = -0.131t + 4.605

-0.975 = -0.131t

\[ \textbf{7.4 d} = t \]
IV. Nuclear Fission / Fusion
(splitting/breaking apart) (combining)

Fission: \[ {}^0_1n + {}^{235}_{92}U \rightarrow {}^{52}_{137}Te + {}^{97}_{40}Zr + 2 {}^0_1n \] (nuclear power plants)
\[ {}^0_1n + {}^{239}_{94}Pu \rightarrow {}^{146}_{56}Ba + {}^{91}_{38}Sr + 3 {}^0_1n \] (nuclear bombs)

Subcritical mass: the reaction dies out.
Critical mass: for each neutron that reacts, one neutron that is produced, causes a reaction.
Supercritical mass: exponential increase, explosion.

Fusion on the sun:
\[ {}^1_1H + {}^1_1H \rightarrow {}^2_2He + {}^0_1e \]
\[ {}^1_1H + {}^2_2H \rightarrow {}^3_2He \]
\[ {}^2_3He + {}^2_3He \rightarrow {}^4_2He + 2 {}^1_1H \]
\[ \text{or} \]
\[ {}^2_3He + {}^1_1H \rightarrow {}^4_2He + {}^0_1e \]

In Fission and Fusion reactions some mass is lost and converted to energy! \( E = mc^2 \)

V. Nuclear Power Plants
1) Fuel: \(_{92}^{235}U \) & \(_{94}^{239}Pu\) (produced in the reaction)

2) Moderator: \( H_2O, \) heavy \( H_2O, \) graphite, Be
   -slows down neutrons, but does not absorb them.

3) Control Rods: B, Cd, or Ga alloyed with Al or steel
   -absorbs excess neutrons not needed to keep the chain reaction going.

4) Coolant: He, \( H_2O, \) CO\(_2\), liquid metals (Na)
   -removes heat from the reactor.

5) Heat Exchanger:
   -transfers heat from the hot coolant, to the cool water, that will make steam to turn the turbines.

6) Safety Shields: Pb or thick concrete
   -radiation shield around the reactors.

*End of Notes* (Assignments #98-99 are Review Assignments. There are no notes for these assignments.)