

1. Complete the following table:

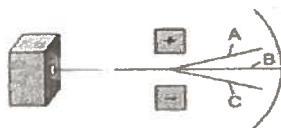
	Isotopic Notation	# of Protons	# of Neutrons	Mass Number	Element Name
a.	${}^2_1\text{H}$	1	1	2	Hydrogen-2
b.	${}^{37}_{17}\text{Cl}$	17	20	37	chlorine-37
c.	${}^{107}_{47}\text{Ag}$	47	60	107	Silver-107

2. Complete the following table:

	Type	Definition	Mass	Charge	Penetrating Ability (energy)
α	Alpha (${}^4_2\text{He}$)	helium nucleus	4	+2	Lowest (stopped by paper/skin)
β^-	Beta (${}^0_{-1}\text{e}$)	high energy electron	0	-1	Medium (stopped by foil)
β^+	Positron (${}^0_{+1}\text{e}$)	+ charged electron	0	+1	
γ	Gamma (γ)	high energy ray	0	0	Highest (stopped by 1m lead/concrete)

3. Radiation is streaming out of the lead box and between a positively and a negatively charged plate.

What type of particle is A beta, B gamma, C alpha
(electron)

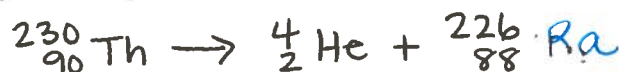


4. Write a nuclear equation for the alpha decay of:

a. Radium-226

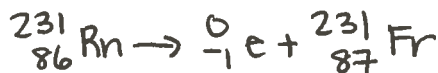


b. Thorium-230



5. Write a nuclear equation for the beta decay of:

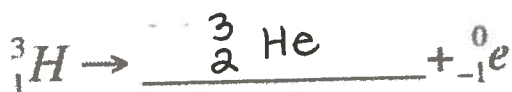
a. Radon-231



b. Krypton-81



6. Balance each nuclear equation. Identify the type of nuclear equation next to it.



beta decay

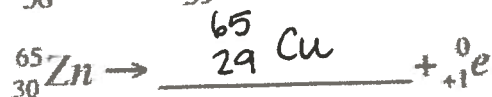
(decay = emission)



alpha decay



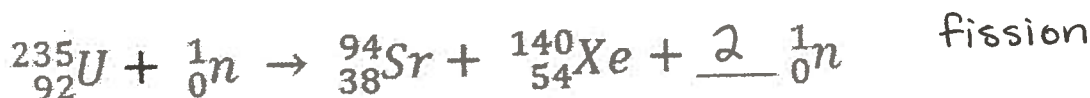
beta decay



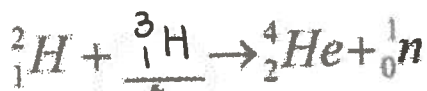
positron decay



positron decay



fission



fusion

7. The half-life of chromium-51 is 28 days. If you have 2.00 grams of chromium, how many grams would remain after 56 days?

$$t_{1/2} = 28 \text{ d}$$

$$A_0 = 2.00 \text{ g}$$

$$t = 56$$

$$A = ?$$

$$n = \frac{t}{t_{1/2}} = \frac{56 \text{ d}}{28 \text{ d}} = 2 \quad \text{so} \quad A = \frac{A_0}{2^n} = \frac{2.00}{2^2} = 0.500 \text{ g}$$

8. In the year 1992, a doctor's office purchased 2 grams of Cesium-137 that registered 400 radiation counts per second (cps). Given the half-life of Cesium-137 is 30 yrs, how much radiation would be expected to remain by the year 2016?

$$2016 - 1992 = 24 \text{ yrs} = t$$

$$t_{1/2} = 30 \text{ yrs}$$

$$A = ?$$

$$A_0 = 2 \text{ g}$$

$$n = \frac{t}{t_{1/2}} = \frac{24}{30} = .8 \text{ yrs}$$

$$\text{so } A = \frac{A_0}{2^n} = \frac{2}{2^{.8}} \approx 1 \text{ g (really 1.1487 g)}$$

9. A 208 g sample of sodium-24 decays to 13.0 g of sodium-24 within 60.0 h. What is the half-life of this radioactive isotope?

$$A_0 = 208 \text{ g}$$

$$A = 13.0 \text{ g}$$

$$t = 60.0 \text{ h}$$

$$t_{1/2} = ?$$

$$A = \frac{A_0}{2^n}$$

$$\text{so } 13.0 = \frac{208}{2^n}$$

$$13.0 (2^n) = 208$$

$$2^n = 16$$

$$n = 4$$

$$\text{and } n = \frac{t}{t_{1/2}} \text{ so } t_{1/2} = \frac{t}{n} = \frac{60}{4} = 15.0 \text{ h}$$

10. The half-life of protactinium-234 is 6.75 hours. What percentage of a given sample will remain after 27 hours?

$$100\% \text{ remains @ } t=0 \text{ so } A_0 = 100$$

$$A = ?$$

$$t_{1/2} = 6.75 \text{ hrs}$$

$$t = 27 \text{ hrs}$$

$$n = \frac{t}{t_{1/2}} = \frac{27}{6.75} = 4, \text{ so } A = \frac{A_0}{2^n} = \frac{100}{2^4} = 6.25\%$$

11. Compare nuclear fusion to nuclear fission, with regards to their definitions, the amount of energy they produce per mole, and their uses.

fusion - in sun

- 2 or more smaller nuclei form 1 larger nucleus
- needs ↑↑ temp + pressure
- releases more e. per mole
- no nuc. waste

fission - in nuc. power plants

- triggered by neutron bombardment
- 1 large nucleus breaks up into smaller nuclei
- releases less e. per mole
- produces nuc. waste

12. When solid barium hydroxide octahydrate is mixed with solid ammonium chloride in a glass flask, the flask gets cold.

- Is the reaction absorbing or releasing heat? *absorbing*
- Is the reaction endothermic or exothermic? *endothermic*
- Why does the flask get cold? *the flask is part of the surroundings and it lost energy to the system (i.e. the system gained energy)*

13. How much heat is released when 620 g of iron (C of iron = $0.444 \text{ J/g} \cdot ^\circ\text{C}$) drops from 80.6°C to 46.2°C ?

$$\begin{aligned} T_i &= 80.6^\circ\text{C} \\ T_f &= 46.2^\circ\text{C} \\ \Delta T &= -34.4^\circ\text{C} \\ m &= 620 \text{ g} \\ C &= .444 \text{ J/g} \cdot ^\circ\text{C} \end{aligned}$$

$$\begin{aligned} q &= mC\Delta T \\ &= 420 \text{ g} \cdot 0.444 \text{ J/g} \cdot ^\circ\text{C} \cdot -34.4^\circ\text{C} \\ &= \boxed{-6400 \text{ J}} \text{ or } \boxed{-6.4 \times 10^3 \text{ J}} \end{aligned}$$

14. A 16.5 g sample of a metal is warmed to 86.5°C and added to 50.0 g of 19.0°C water. The final temperature of the water is 20.7°C . What is the specific heat of the metal? *The metal and water have the same*

$$\begin{aligned} m_{\text{met.}} &= 16.5 \text{ g} & m_{\text{H}_2\text{O}} &= 50.0 \text{ g} \\ T_{i \text{ metal}} &= 86.5^\circ\text{C} & T_{i \text{ H}_2\text{O}} &= 19.0^\circ\text{C} \\ T_{f \text{ metal}} &= 20.7^\circ\text{C} & T_{f \text{ H}_2\text{O}} &= 20.7^\circ\text{C} \end{aligned}$$

$$T_f (= 20.7^\circ\text{C})$$

$$\begin{aligned} C_{\text{metal}} &= \frac{50.0 \cdot 4.184 \cdot (20.7 - 19.0)}{-(16.5 \cdot (20.7 - 86.5))} \\ &= \boxed{0.328 \text{ J/g} \cdot ^\circ\text{C}} \end{aligned}$$

$$\begin{aligned} -q_{\text{metal}} &= q_{\text{H}_2\text{O}} \\ -(m_{\text{met.}} \cdot C_{\text{met.}} \cdot \Delta T_{\text{met.}}) &= (m_{\text{H}_2\text{O}} \cdot C_{\text{H}_2\text{O}} \cdot \Delta T_{\text{H}_2\text{O}}) \\ C_{\text{metal}} &= \frac{m_{\text{H}_2\text{O}} \cdot C_{\text{H}_2\text{O}} \cdot \Delta T_{\text{H}_2\text{O}}}{-(m_{\text{metal}} \cdot \Delta T_{\text{metal}})} \end{aligned}$$

15. What mass of 58.0°C iron must be added to 342 g of 23°C ethanol (C of ethanol = $2.46 \text{ J/g} \cdot ^\circ\text{C}$) to make the final temperature of both come out to be 40.0°C ? *T_f is the same for iron and ethanol*

$$\begin{aligned} m_{\text{Fe}} &=? & m_{\text{etha.}} &= 342 \text{ g} \\ C_{\text{Fe}} &= 0.444 \text{ J/g} \cdot ^\circ\text{C} & C_{\text{etha.}} &= 2.46 \text{ J/g} \cdot ^\circ\text{C} \\ T_i &= 58.0^\circ\text{C} & T_i &= 23^\circ\text{C} \\ T_f &= 40.0^\circ\text{C} & T_f &= 40.0^\circ\text{C} \end{aligned}$$

$$\begin{aligned} -q_{\text{Fe}} &= q_{\text{ethanol}} \\ -(m_{\text{Fe}} \cdot C_{\text{Fe}} \cdot \Delta T_{\text{Fe}}) &= (m_{\text{etha.}} \cdot C_{\text{etha.}} \cdot \Delta T_{\text{etha.}}) \\ m_{\text{Fe}} &= \frac{m_{\text{etha.}} \cdot C_{\text{etha.}} \cdot \Delta T_{\text{etha.}}}{-(C_{\text{Fe}} \cdot \Delta T_{\text{Fe}})} = \frac{342 \cdot 2.46 \cdot (40.0 - 23)}{-(0.444 \cdot (40.0 - 58.0))} \\ m_{\text{Fe}} &= \boxed{1790 \text{ g}} \end{aligned}$$

16. List the endothermic phase changes: *melting, vaporization, sublimation*

17. List the exothermic phase changes: *freezing, condensation, deposition*

18. During a phase change why does the temperature stay constant? *Average kinetic energy is constant during phase changes. Only potential energy changes as attractions between molecules change.*

Use the heating curve of water and the H_{fus} , H_{vap} and C values in your packet to answer questions #19 – 25.

19. Between what two points does melting (fusion) occur for this substance?

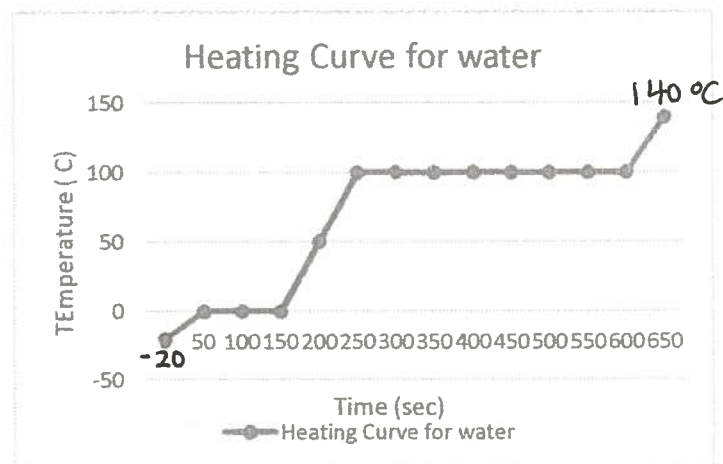
$$t = 50 \text{ to } t = 150$$

20. Between what two points is this substance in the solid phase? The liquid phase? The gas phase?

Solid: $t = 25 \text{ to } t = 50$ gas: $t = 600 \text{ to } t = 650$
 liquid: $t = 150 \text{ to } t = 250$

21. Between what two points could the heat of vaporization for this substance measured?

$$t = 250 \text{ and } t = 600$$



22. Which line segments on the curve show an increase in average kinetic energy? How about an increase in potential energy? Inc. in avg. KE: $t = 25 \text{ to } t = 50$, $t = 150 \text{ to } t = 250$, $t = 600 \text{ to } t = 650$
 Inc. in potential energy: $t = 50 \text{ to } t = 150$, $t = 250 \text{ to } t = 600$

23. How much heat was required to heat 50.0 grams of water from the lowest temperature on the graph to the highest temperature on the graph?

$$q = m C_s \Delta T$$

$$= 50.0 \text{ g} \cdot 2.092 \text{ J/g} \cdot ^\circ\text{C} \cdot (20^\circ\text{C})$$

$$= 2092 \text{ J}$$

$$q = m H_{fus}$$

$$= 50.0 \text{ g} \cdot 334 \text{ J/g}$$

$$= 16700 \text{ J}$$

$$q = m C_l \Delta T$$

$$= 50.0 \cdot 4.184 \cdot (100^\circ\text{C})$$

$$= 20920 \text{ J}$$

$$q = m H_{vap}$$

$$= 50.0 \cdot 2260 \text{ J/g}$$

$$= 113000 \text{ J}$$

$$q = m C_g \Delta T$$

$$= 50.0 \text{ g} \cdot 2.006 \text{ J/g} \cdot ^\circ\text{C} \cdot 40^\circ\text{C}$$

$$= 4012 \text{ J}$$

$$\text{total (add)} = 453395712 \text{ J}$$

$$= 4.53 \times 10^8 \text{ J}$$

24. How much heat in joules does it take to vaporize 0.5000 L of H_2O at 100°C ? ($D_{\text{water}} = 1 \text{ g/ml}$) $H_{vap} = 2260 \text{ J/g}$

$$q = m H_{vap}$$

$$= 500.0 \text{ g} \cdot 2260 \text{ J/g} = 1.130 \times 10^6 \text{ J}$$

25. How much heat in kilojoules is released when 12.90 mL of H_2O freezes at 0°C ? $H_{fus} = 334 \text{ J/g}$

$$q = m H_{fus}$$

$$= 12.90 \text{ g} \cdot 334 \text{ J/g} = 4309 \text{ J} = 4.309 \text{ kJ}$$

$$(1 \text{ kJ} = 1000 \text{ J})$$

26. (Lab) For the heat of fusion of ice lab...what errors or occurrences would cause the calculated heat of fusion to be larger than it should be? Smaller than it should be? How do they affect the calculation of the heat of fusion?

$$H_{fus} = \frac{q_{\text{melted ice}}}{m_{\text{melted ice}}} = \frac{-q_{\text{tap water}}}{m_{\text{melted ice}}} = \frac{-(m_{\text{tap}} \cdot 4.184 \cdot \Delta T_{\text{tap}})}{m_{\text{melted ice}}}$$

Larger H_{fus} :

- Reading the meniscus of the tap water too high (m_{tap} too high) so $H \uparrow$
- Spilling melted ice water while removing ice chunks ($m_{\text{melted ice}}$ too low) so $H \uparrow$

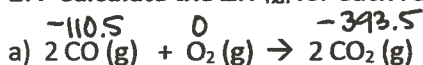
Smaller H_{fus} :

- Not drying ice cubes before adding to cup ($m_{\text{melted ice}}$ too large) so $H_{fus} \downarrow$
- Reading T_f on the graph too high (ΔT_{tap} too small) so $H_{fus} \downarrow$

$$\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_{\text{prod}} - \sum \Delta H^\circ_{\text{reactants}}$$

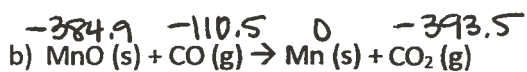
For questions #27-28, use this chart.

27. Calculate the $\Delta H^\circ_{\text{rxn}}$ for each reaction:



$$\Delta H^\circ_{\text{rxn}} = [2(-393.5)] - [2(-110.5)]$$

$$= -566.0 \text{ kJ}$$

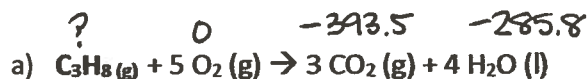


$$\Delta H^\circ_{\text{rxn}} = [0 + (-393.5)] - [-384.9 + (-110.5)]$$

$$= +101.9 \text{ kJ}$$

28. Calculate the ΔH°_f of the chemical in bold given the

ΔH° :



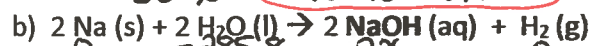
$$\Delta H^\circ_{\text{comb}} = -2220.1 \text{ kJ}$$

$$-2220.1 \text{ kJ} = [3(-393.5) + 4(-285.8)] - x$$

$$-2220.1 = -2323.7 - x$$

$$103.6 = -x$$

$$\text{so } x = -103.6 \text{ kJ/mol}$$



$$\Delta H^\circ_{\text{rxn}} = -184 \text{ kJ}$$

$$-184 \text{ kJ} = [2x + 0] - [2(0) + 2(-285.8)]$$

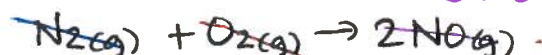
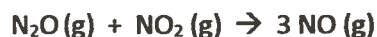
$$-184 = 2x - (-571.6)$$

$$-755.6 = 2x \quad \text{so } x = \frac{-755.6}{2} = -377.8 \text{ kJ/mol}$$

29. Given the following data:



Use Hess's law to calculate ΔH_{rxn} for the following reaction.



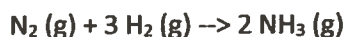
$$\Delta H = 180.7$$



$$\Delta H = +156.1 \text{ kJ}$$

Compound	ΔH_f° (kJ/mol)	Compound	ΔH_f° (kJ/mol)
$\text{CH}_4(\text{g})$	-74.8	HCl(g)	-92.3
$\text{CO}_2(\text{g})$	-393.5	$\text{H}_2\text{O(g)}$	-241.8
CO(g)	-110.5	$\text{SO}_2(\text{g})$	-296.1
$\text{H}_2\text{O(l)}$	-285.8	$\text{NH}_4\text{Cl(s)}$	-315.4
$\text{H}_2\text{S(g)}$	-20.1	NO(g)	+90.4
$\text{H}_2\text{SO}_4(\text{l})$	-811.3	$\text{NO}_2(\text{g})$	+33.9
$\text{MgSO}_4(\text{s})$	-1278.2	$\text{SnCl}_4(\text{l})$	-545.2
MnO(s)	-384.9	SnO(s)	-286.2
$\text{MnO}_2(\text{s})$	-519.7	$\text{SnO}_2(\text{s})$	-580.7
NaCl(s)	-411.0	$\text{SO}_3(\text{g})$	-296.1
NaF(s)	-569.0	$\text{SO}_3(\text{l})$	-395.2
NaOH(s)	-426.7	ZnO(s)	-348.0
$\text{NH}_3(\text{g})$	-46.2	ZnS(s)	-202.9

30. Calculate ΔH for the reaction $\text{CH}_4(\text{g}) + \text{NH}_3(\text{g}) \rightarrow \text{HCN}(\text{g}) + 3 \text{H}_2(\text{g})$, given:



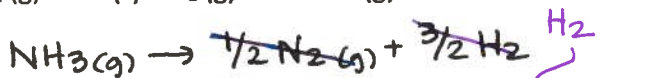
$$\Delta H = -91.8 \text{ kJ} \quad \text{flip, } \div \text{ by } 2$$



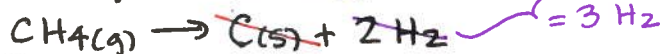
$$\Delta H = -74.9 \text{ kJ} \quad \text{flip}$$



$$\Delta H = +270.3 \text{ kJ} \quad \div 2$$



$$\Delta H = +91.8/2$$



$$+ \Delta H = +74.9$$

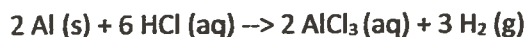


$$+ \Delta H = 270.3/2$$

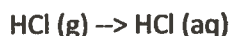


$$\Delta H = 256.0 \text{ kJ}$$

31. Calculate ΔH for the reaction $2 \text{Al}(\text{s}) + 3 \text{Cl}_2(\text{g}) \rightarrow 2 \text{AlCl}_3(\text{s})$ from the data.



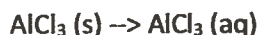
$$\Delta H = -1049. \text{ kJ} \quad \text{keep}$$



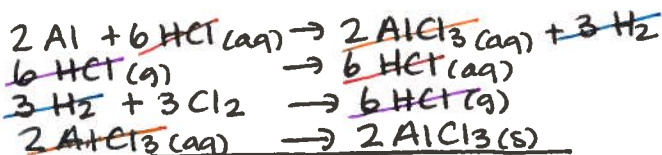
$$\Delta H = -74.8 \text{ kJ} \quad \times 6$$



$$\Delta H = -1845. \text{ kJ} \quad \times 3$$



$$\Delta H = -323. \text{ kJ} \quad \text{flip, } \times 2$$



$$\Delta H = -1049 \text{ kJ}$$

$$\Delta H = -74.8 \times 6$$

$$\Delta H = -1845 \times 3$$

$$\Delta H = +323 \times 2$$

$$\text{total } \Delta H = -7679 \text{ kJ}$$

32. (Lab) The following data chart is generated for a food calorimetry lab where a piece of food was burned and the generated heat was released to a "soda can" type of calorimeter.

Mass of water (g)	50.0 g
Initial temperature of water (degrees C)	18.8 degrees C
Final temperature of water (degrees C)	39.7 degrees C

a) Calculate the heat absorbed by the water. $q = mC\Delta T = 50.0 \text{ g} \cdot 4.184 \cdot (39.7 - 18.8) = 4370 \text{ J}$

b) How many joules of energy are contained in the burned food? 4370 J (assuming all heat entered the calorimeter)

c) How many heat calories and dietary Calories are in the burned food?

$$\frac{4370 \text{ J}}{4.184 \text{ J}} = 1045 \text{ cal (heat cal.)} \quad \text{and} \quad \frac{1045 \text{ cal}}{1000 \text{ cal}} = 1.045 \text{ Calories}$$

d) What effect does each of these occurrences have on the calculated energy content of the food? Explain.

- Reading the final temperature of water as lower than the true value. ΔT low so ΔT will be too low q too low so #Cal too low
- Spilling water from the can after its volume was measured and before the experiment. $m_{\text{H}_2\text{O}}$ too low so q will be too low \therefore #Cal too low
- There is open space between the burning food and the calorimeter. q will be too low (too much heat escapes) so #Cal too low